
LABORATORY AND
WORKBOOK UNITS

IN

CHEMISTRY

BY

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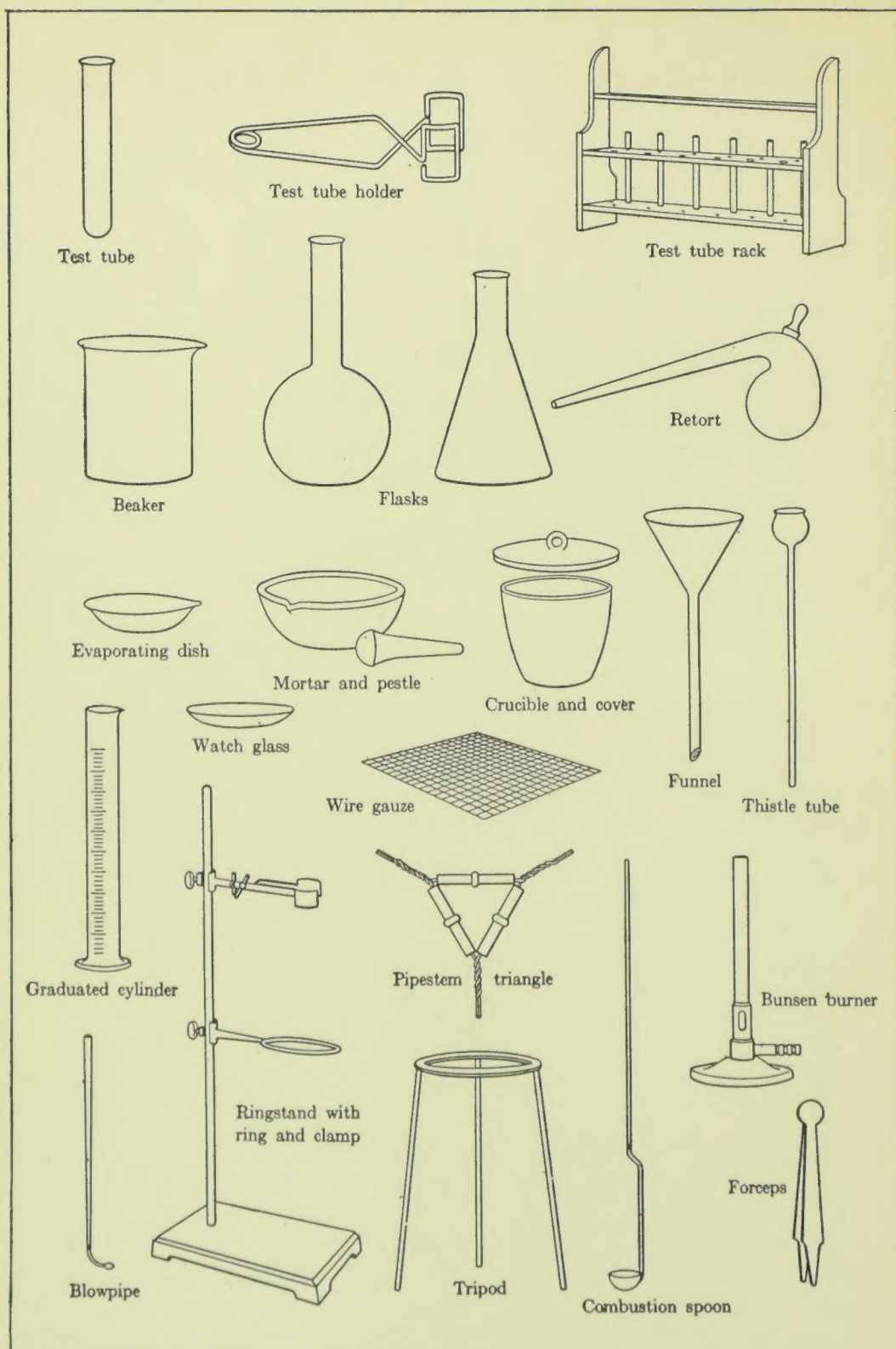
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Laboratory apparatus in common use. You should learn to recognize each of these pieces of apparatus and to be able to find them quickly in the laboratory.

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IN
CHEMISTRY

BY

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TO THE TEACHER

Chemistry is an experimental science. It is through experimental work in the school chemistry laboratory that students make simple chemical discoveries and applications. Through experimental verification, they are convinced of the truth of the laws and theories of chemistry. But unless this work is related to and interwoven into classroom work, much of the educational value of experimental work is lost. The student needs guidance in the selection, utilization, and application of what he has learned in the laboratory. Demonstrations witnessed by the student, individual laboratory work, classroom work with the textbook, and reference work in the library must be integrated.

The authors hope that these purposes have been so well incorporated in **LABORATORY AND WORKBOOK UNITS IN CHEMISTRY** that the student will be aided in realizing the objectives of a study of chemistry.

The fifty-seven units of this book are designed to meet the needs of a complete course in elementary chemistry. They are based on the suggestions and requirements of state and college entrance syllabi, on standard textbooks, and on many years experience on the part of the authors as teachers of elementary chemistry.

Each unit begins with **EXPERIMENTS**. Here are given detailed instructions for the performance of experimental work. Such questions as "Result?," "Color?," and the like are included to arouse thought and stimulate observation.

The **EXPERIMENTS** are followed by **OBSERVATIONS AND QUESTIONS ON THE EXPERIMENTS**. Here the student is led to record his observations in the laboratory. Questions that are worded objectively help him to organize and report what he has seen and done. In the next section of the unit called **CONCLUSIONS**, he is led to the formation of generalizations based upon his observations. In this way he arrives at scientific conclusions concerning the basic facts taught in the unit.

SUPPLEMENTARY EXERCISES follow. These correlate the laboratory work with the work of the text. The purpose of these exercises is to arouse the student's scientific curiosity and interest. They will cause him to spend a maximum of time in thinking and in constructive work and a minimum of time in writing.

OPTIONAL QUESTIONS of the essay type are given at the close of each unit in order to give practice and experience with this type of question. They include a few questions which will tax the ability of even the most gifted student. Although most of the questions and exercises in both SUPPLEMENTARY EXERCISES and OPTIONAL QUESTIONS may be answered by the use of a standard text, a few have been included which will demand work with reference books.

Clearly labeled diagrams of laboratory setups and pictures of standard laboratory apparatus have been placed at the beginning of certain units to help the student in arranging and setting up the apparatus for use in his experimental work. These diagrams also will serve as models for the ones to be drawn by the student from his own laboratory setups.

The lists of apparatus and materials, customarily placed at the head of each experiment, are placed instead at the back of the book for convenient reference by the teacher, the laboratory assistant, or the student. The experiments have been so planned that they may be performed with simple apparatus and inexpensive materials in small amounts.

The factors—time, physical equipment, cost of supplies, type of student, syllabi requirements—will suggest to each teacher which of the experiments listed should be demonstrations, which should be performed as individual laboratory work, and which should be omitted. For some experiments the demonstration method must be used because of the need for delicate and expert technique or because of their dangerous nature. However, there are undoubtedly many experiments best and most profitably performed individually by the student in the laboratory. The authors have suggested which experiments might be demonstrated, but obviously

each teacher must decide this question for himself upon the basis of his individual situation.

LABORATORY AND WORKBOOK UNITS IN CHEMISTRY has been planned for use either in schools which have a single laboratory period or in those which have a double laboratory period. Practical experience in the laboratory has shown that the section of each unit called EXPERIMENTS can be completed in a single period.

As is evident from the nature of the material in each unit, LABORATORY AND WORKBOOK UNITS IN CHEMISTRY makes provision for individual differences. The SUPPLEMENTARY EXERCISES and OPTIONAL QUESTIONS enable the teacher to adjust the course to meet the individual needs of students. Thus each student may be kept working at a maximum efficiency.

The workbook type of laboratory manual is well adapted to lead students in recording important observations made in the laboratory and in drawing correct conclusions from these observations. It is for this reason that the workbook type of laboratory manual has met such widespread approval from teachers of science. But in many schools, consumable books which must be replaced each year are out of the questions because of cost, school regulations, and the like. It is for such schools that LABORATORY AND WORKBOOK UNITS IN CHEMISTRY, Non-consumable Edition, has been planned. In this edition, is presented a laboratory manual and workbook in chemistry which retains the essence of the workbook but which is written in non-consumable form.

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TO THE STUDENT

Not until ancient alchemy had reached the stage when men started to test their beliefs in the laboratory did alchemy give way to modern chemistry. It became a real science when men forgot their search for the *philosopher's stone* which would transform other metals into gold and gave up their vain attempts to concoct an *elixir of life* which would ward off old age. When men turned from such fantastic goals to simple experiments, they began to discover the laws of chemistry. Thus chemistry was born in the laboratory.

In 1774, Priestley, an amateur chemist, heated a red powder (mercuric oxide) and obtained from it the gas, which we call oxygen. He experimented with this gas in his kitchen laboratory and discovered that in it a candle burned with great splendor. He described his experiments to the French scientist, Lavoisier, whose scientific imagination led him to think that in this gas might be found an explanation of an age-old puzzle—an explanation of burning. Lavoisier went to his laboratory and there heated mercury in a sealed flask until all of it had changed into the red powder used by Priestley. He continued to heat this red powder and from it obtained oxygen just as Priestley had done. By weighing all of the materials used in his experiments with the most sensitive balance in all France, Lavoisier found incontrovertible evidence that burning is nothing more than the rapid union of a burning substance with oxygen. This generalization made by Lavoisier as a result of his experiments started chemistry on the glorious road which has led to the compounds, alloys, and medicines that have helped to transform modern life.

The great pioneers of chemistry of a century and a half ago tore themselves away from the teachings of the ancients and trusted only in the facts born of their experiments in the laboratory. The student of chemistry today is given the opportunity of repeating many of these classic experiments and from them discovering for himself the truths which the great pioneers established. Experiments in the laboratory and demonstrations by the

instructor transform the cold, inanimate facts of chemistry into living concepts. Such experimental work will give you the thrill of discovery and verification and will help you to understand how the chemist lives and works.

Laboratory work in high school should be an awakener and a clarifier, and it is hoped that for some students it may be preparation for chemical research. Some of you may, in the future, make as great discoveries as any of the pioneers of chemistry. These great pioneers did their laboratory work carefully, systematically, and honestly. Their method of work should be an inspiration to you. Such work will involve preparation before you come to the laboratory. It will mean also keeping a neat, carefully-written record of your work. In your laboratory work, you will be adventuring into the history of the development of chemistry and of its multifold applications in the everyday world.

As the result of years of experience by thousands of chemists, certain procedures have been established which the successful worker in the chemical laboratory observes. You should observe these procedures carefully and honestly not only in the chemical laboratory but in every situation in which laboratory technique is used.

1. Understand the meaning and purpose of all directions for each experiment before you attempt to perform it. Laboratory work should be thoughtful, not mechanical or formal.

2. As you perform an experiment, observe the changes that are taking place. Try to understand the meaning of each change. These changes may be changes in color, the giving off of a gas, or the like.

3. Do not waste chemicals or use more than necessary. Reactions usually take place faster and more efficiently between small amounts of chemicals. Put all waste solids into the earthenware crocks provided expressly for that purpose. Liquids may be poured into the sink, but should be flushed with plenty of water.

4. Use only the chemicals and the quantities indicated. By handling chemicals and apparatus carefully and thoughtfully, you

will avoid accidents and injuries. **In case of an accident, report to your instructor immediately.**

To make the most efficient use of this book, you should, upon completion of the experiments in a unit, enter your observations and conclusions in your notebook by copying and completing the OBSERVATIONS AND QUESTIONS ON EXPERIMENTS and CONCLUSIONS. Draw neat, accurate diagrams of the apparatus used.

You should then copy and complete as many of the SUPPLEMENTARY EXERCISES and answer as many of the OPTIONAL QUESTIONS as time permits. You will need to use your textbook in this work. A few of the questions and exercises may require reference work in the school library.

We hope you will enjoy your laboratory work and that, in using this book, you will acquire the scientific attitude of testing and proving ideas presented to you instead of accepting them on faith.

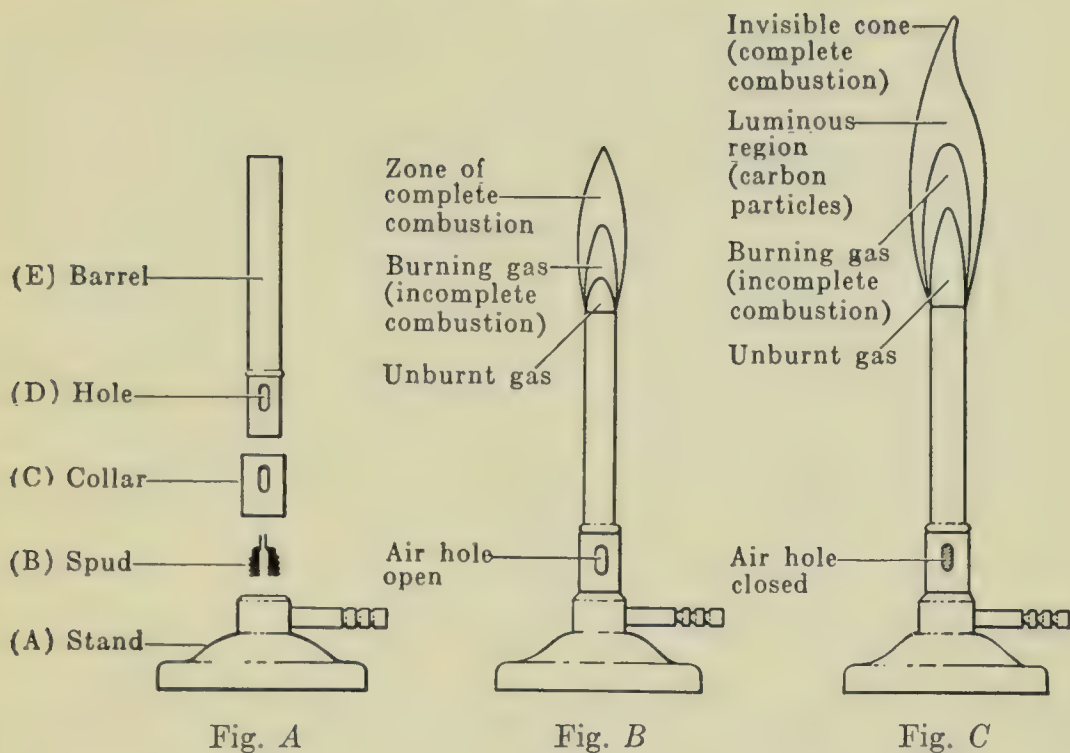
THE AUTHORS

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Preliminary Experiment with a Bunsen Burner and Glass Tubing



1. Take a Bunsen burner apart and locate each of the parts shown in Fig. A above.

2. Reassemble the Bunsen burner. Attach it with a piece of rubber tubing to the source of the fuel gas. Turn on the gas and ignite the burner by holding a match at the top of the barrel.

3. By turning the collar, open the air hole. What kind of flame do you obtain? This is a nonluminous flame. It is shown in Fig. B above. Locate each of the three cones in the nonluminous flame.

4. By turning the collar, shut the air hole on the Bunsen burner. What kind of flame do you obtain? This is a luminous flame. How does the luminous flame differ from the nonluminous flame? Find each of the three cones in the luminous flame. (See Fig. C.)

5. Hold a cold, dry test tube in a luminous flame for about 30 seconds. Is there a deposit? Hold a second cold, dry test tube in a

nonluminous flame for about 30 seconds. Is there a deposit? Explain the difference in the results.

NOTE. A nonluminous flame with a "roaring" sound indicates that too much air is entering the burner. This can be corrected by closing the air hole slightly. If the flame "strikes back" and burns noisily at the spud, it means that there is insufficient gas pressure and too much air. If the flame should "strike back," turn off the gas and start over. Be sure to give the burner a chance to cool before touching it again.

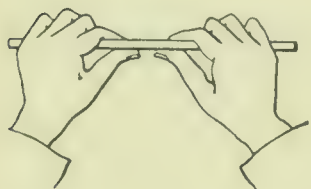


Fig. D

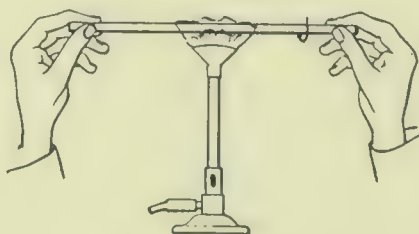


Fig. E

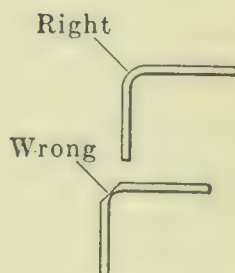


Fig. F

6. With a triangular file make a deep scratch in the middle of a 12-inch piece of glass tubing. Hold the tubing as shown in Fig. D above. Pressing forward with your thumbs, break the tubing in half.

7. Glass tubing which has just been broken has sharp edges. "Fire polish" these edges by holding the end of the rough tube in the upper part of a nonluminous flame, rotating the tube slowly, until the edge has been made smooth by the melting of the glass.

8. Turn off the gas. Place the fishtail tip on the burner as shown in Fig. E. Now rotate one of the 6-inch pieces of glass in the nonluminous flame as shown. When the flame becomes yellow in color and the softened glass begins to bend of its own weight, remove the glass and bend it into a right angle, as in Fig. F. Notice that the bend should be smooth and round.

9. Hot glass, if cooled too quickly, will become brittle. To avoid this, glass is annealed (cooled gradually). Anneal the right angle bend you made by slowly rotating it in a luminous Bunsen flame until the bent portion is completely covered with a deposit of soot. Allow the glass to cool; then remove the soot with a rag.

INTRODUCTORY UNIT. HOW WE LEARN, THINK AND SOLVE PROBLEMS

Modern life and modern living conditions are vastly different today than they were one hundred years ago even though one hundred years in the history of man is a very short time. Many of the changes which brought about modern civilization are direct results of the work of scientists who, insofar as the material things of life are concerned, are continuously changing our ways of living.

How have scientists produced these amazing changes? It is generally agreed that their *methods* of learning, thinking, and solving problems are largely responsible for the rapid forward strides of science. We should try to master these methods so that we may use them in helping to solve our everyday economic, social, and political problems.

In the material which follows, do not fill in the blanks in this book. Write only in your notebook.

1. WE LEARN BY OBSERVATION

(a) Most of us have looked at a bottle of peroxide of hydrogen or tincture of iodine or catsup many times either at home or elsewhere. Try to remember what was printed on the label of one of these bottles.

List the items that you remember having observed on this label.

- | | |
|----------|----------|
| 1. _____ | 3. _____ |
| 2. _____ | 4. _____ |

(b) Now check your memory of your observations by actually examining this label. List the items on the label by direct observation.

- | | |
|----------|----------|
| 1. _____ | 4. _____ |
| 2. _____ | 5. _____ |
| 3. _____ | 6. _____ |

(c) Can you trust completely your memory of observable facts of this kind? How might you help yourself to remember facts or

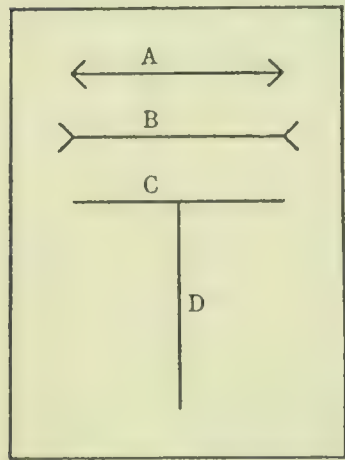
observations that seem worth retaining?-----

(d) The senses by means of which we observe are: sight,
-----, -----, -----, and -----.

(e) By what means other than the label could you identify a
bottle of peroxide of hydrogen or tincture of iodine or catsup?

1. -----
2. -----
3. -----
4. -----

(f) Are we always sure that things really are as they appear
to us? Is line *A* the same length as line
B? Is line *C* the same length as line
D? Test your observation by placing a ruler
or other straightedge along these lines.
Were your first observations correct?
-----.



(g) What are the defects of, and what
safe-guards must be used in, learning by
observation? -----

2. WE LEARN BY ASKING QUESTIONS

(a) We ask questions of teachers, neighbors, parents, informa-
tion bureaus, doctors, lawyers, etc. We ask questions of diction-
aries, encyclopedias, and books of various kinds. Such people and
such sources are sometimes called *experts* or *authorities*.

What objections can you think of against acquiring factual
information in this way? -----

What advantages can you think of for acquiring factual infor-
mation in this way? -----

(b) Would you like an expert or an authority in a given field such as your teachers or textbooks, to solve all your problems for you? ----- Would you like to solve some problems yourself? ----- Why? -----

(c) Would you go to the same expert to get help in answering the following questions? "How would I cure a skin rash?" "How can I remove a particular stain from my clothing?" "How can I get my radio to work?" Why? -----

(d) Why do many of us feel that we must go to an expert to get help in solving many of the problems that confront us?

3. WE LEARN BY DOING

(a) Can you drive an automobile? Can you play tennis well? Can you bend a piece of glass tubing in the shape of an "L"? How would you proceed in learning these skills? Would reading or asking about them be enough? -----

(b) Suppose you were beginning the study of oxygen. Could you learn about oxygen by referring to your textbook? By watching your teacher prepare oxygen and perform some experiments using it? By preparing oxygen and performing some experiments using it yourself? Which method, or methods, would you use? Why? -----

4. WE LEARN BY REASONING

(a) Aristotle and other thinkers of ancient times believed that one of the best methods of obtaining knowledge was by reasoning. They paid little or no attention to observation. Aristotle reasoned that heavy objects fall faster than lighter ones. By performing an actual experiment and observing what happened, Galileo proved that Aristotle was wrong. Why was Aristotle's conclusion incorrect? -----

(b) Lavoisier, the “Father of Modern Chemistry,” *burned* tin in air. He observed that the appearance of the tin had changed and that it increased in weight. What conclusion would you draw by reasoning from these two facts? -----

(c) Lavoisier *burned* mercury in air and observed that it, also, increased in weight. A red powder was formed. He heated this red powder strongly and obtained mercury and a gas as products. What conclusion would you draw by reasoning from these facts? -----

(d) In learning by reasoning, what should we try to be sure of? -----

5. WE SOLVE PROBLEMS BY MEANS OF THE EXPERIMENTAL METHOD OF SCIENCE

List the six steps, or phases, of the method of science. Opposite each step write the corresponding step followed by Lavoisier in answering the question, “What happens when a substance burns in air?”

Steps in the Scientific Method.	Steps Lavoisier Used.
1. ----- -----	1. ----- -----
2. ----- -----	2. ----- -----
3. ----- -----	3. ----- -----
4. ----- -----	4. ----- -----
5. ----- -----	5. ----- -----
6. ----- -----	6. ----- -----

Unfortunately, many people observe poorly and thereby obtain *inaccurate* information, or data. Too often they believe what they are told or what they read without testing or checking these opinions or these conclusions with known *accurate* facts. Too often they fail to reason clearly or effectively. In the study of chemistry, you will learn to use the method of science in solving not only the problems of science, but also other problems that confront you outside the classroom.

6. PRELIMINARY EXPERIMENTS: LEARNING THE METHOD OF SCIENCE

The teacher lights a candle (a rather large one should be used) and places it in full view of the class. He then blows out the candle flame.

QUESTION (by teacher): "Why does the flame go out?"

Answers (from class):

1. -----
2. -----
3. -----
4. -----
5. -----
6. -----

QUESTION: "How many students feel sure they know why the flame goes out?"

RESULT: ----- students out of ----- students in the class.

The teacher makes no comment, but performs the following experiments, asking each member of the class to observe closely:

1. He claps his hands just above the flame thereby putting it out.
2. He brings a wire gauze down through the flame to the wick and holds it there, thus putting out the flame.
3. He brings a piece of wood or stone down through the flame to the wick and holds it there, thus putting out the flame.
4. He pours water over the flame thus putting out the flame.

QUESTION: "How many students now feel sure they know why the flame was put out?"

RESULT: ----- students out of ----- students in the class.

QUESTION: "Which answer, or answers, that were suggested before shall we accept?"

Answers:

1. -----
2. -----
3. -----

The teacher makes no comment, but performs the following experiments, asking each member of the class to observe closely:

1. He inserts a narrow glass jet tube into the center of the flame near the wick and ignites the vapor issuing from the jet tube.
2. He blows out the candle flame and ignites the vapor issuing from the wick—but very pointedly ignites the vapor at some distance away from the wick and *does not touch* the wick with the match flame. The flame travels back to the wick. This experiment is repeated once or twice so that it is observed clearly by each member of the class.

QUESTION: How many students now feel sure they know the answer to the original question?

RESULT: ----- students out of ----- in the class.

The teacher makes no comment but performs the following experiment asking each member of the class to observe closely:

In a metal pan, he starts a smoldering fire of some material such as leaves or twigs, which requires a lot of blowing in order to make it burst into a blaze.

The teacher then comments: "I blew on the candle flame thereby putting it out. I blow on this fire and it burns more vigorously."

QUESTION: "How many students now feel sure they know the answer to the original question?"

RESULT: ----- students out of ----- in the class.

QUESTION: "Why are so many of you uncertain and not sure that you know the correct answer to the original question?"

Answer: "Because we don't know enough about *burning*."

Teacher: "Before you can solve a problem or answer a question of this kind to your own satisfaction, you must be quite sure of the facts. I notice that many of you are sorry you 'jumped to conclusions' and ventured opinions which were based on too small a number of observations and thereby on insufficient data. Can you be certain of the accuracy of your answer to a question, or of the accuracy of a conclusion if you do not know the true facts? How can you find out what are the facts with respect to this problem or some similar problems in science?"

Answer: -----

Teacher: "There are many social, political and economic questions which we must solve. What must we be sure of before we attempt to answer them?"

Answer: -----

Teacher: "We seem to be agreed that in order to answer questions on *burning* we should know more of the *facts of burning*. Before any substance can burn, three essential conditions must be fulfilled. If this is accurate, how could we keep a substance from burning?"

Answer: -----

Teacher: "Well, then, let us learn more of the facts of burning and then I am sure everyone in the class will have sufficient data upon which to base a conclusion which he will be confident is accurate and which will not embarrass him later as being a 'snap judgment' or a 'bad guess.' "

UNIT 1. PHYSICAL AND CHEMICAL CHANGES

Experiments

1. Compare the appearance (properties) of the following substances *before* and *after* heating in the Bunsen flame. Make a table like the one below and record your results in it.

(a) A glass rod, (b) sugar (heat a small amount in a dry test tube), (c) copper wire (hold it with a pair of forceps), (d) platinum wire (one end sealed in a glass rod), (e) magnesium ribbon (hold with forceps).

SUBSTANCE HEATED	PROPERTIES BEFORE HEATING	PROPERTIES AFTER HEATING
glass	brittle, clear	malleable, soft
sugar	white powder	black solid
copper	shiny metal	changes color to black
platinum	shiny metal	no change
magnesium	shiny solid	burns brightly, white solid

2. Fill a test tube to a depth of $\frac{1}{4}$ inch with granulated sugar. Add a half test-tubeful of cold water. Place your thumb over the mouth of the test tube and shake vigorously. The resulting clear mixture is called a solution. Taste the solution. Place a small portion of this solution on a watch glass and set aside to evaporate in a warm place. Examine and taste the residue on the watch glass.

Caution: Be careful not to get any sulfuric acid on your skin or clothing.

3. Place about a tablespoonful of granulated sugar in an evaporating dish. Pour in enough concentrated sulfuric acid to cover the sugar. Observe any action in the dish during the next few minutes.

Copy the following in your notebook, supplying the word or words needed to complete each statement. Underline the words you supply.

Observations and Questions on Experiments

1. Four of the characteristics (properties) by means of which we recognize substances are —?—.

2. Heating did not change the identifying properties of the following substances: —?—.

3. Heating the following substances caused the formation of new substances with new properties: —?—.

4. The outstanding property by which we recognize sugar is —?—. When mixed with (dissolved in) water, sugar (does, does not) —?— retain this identifying property. When dissolved in water sugar (changes, does not change) —?— to a new substance.

5. When sugar is mixed with concentrated sulfuric acid, the sugar —?—. This indicates a (physical, chemical) —?— change because the sugar —?—.

Conclusions

1. A chemical change causes the formation of —?— substances with —?— properties.

2. A physical change does not cause —?—.

3. When a substance undergoes a chemical change, its composition (changes, does not change) —?—.

4. The solution of a substance in water is usually a (physical, chemical) —?— change.

Supplementary Exercises

1. Copy each of the following in your notebook and after each write the words *physical change* or *chemical change* to indicate the change involved.

(a) Burning of coal

(b) Breaking of glass

(c) Exploding gasoline in an automobile

(d) Separating cream from milk

(e) The change of water to steam

(f) The change of ice to water

(g) Decaying of fruit

(h) Passing electricity through an electric light bulb

2. Copy each of the following in your notebook and indicate alongside each an identifying property.

(a) Salt

(c) Silk

(e) Wood

(b) Gasoline

(d) Glass

(f) Gold

Optional Questions

1. Explain what is meant by a "physical property" of a substance. Cite illustrations.

2. Explain what is meant by a "chemical property" of a substance. Cite illustrations.

3. How does modern chemistry differ from ancient chemistry (alchemy)?

4. Describe an industry in which chemical knowledge plays an important part.

5. Name three of the pioneers of modern chemistry and tell what each contributed.

UNIT 2. ELEMENTS, COMPOUNDS, AND MIXTURES

Experiments

Caution: Carbon disulfide is very inflammable. Keep it away from flames.

1. Place some powdered sulfur on a sheet of paper. Examine, noting its properties. Touch it with a magnet. Any result? Drop a bit of sulfur into $\frac{1}{6}$ of a test-tubeful of carbon disulfide. Shake thoroughly. Any evidence of solution?

2. Repeat Exp. 1 using iron filings instead of sulfur. Results?

3. Mix thoroughly on a piece of paper about $\frac{1}{2}$ of a test-tubeful (15 g.) of powdered sulfur and about $\frac{1}{4}$ of a test-tubeful (17 g.) of iron filings. Repeat Exp. 1 with some of this mixture. Results?

4. Pour the balance of the mixture of sulfur and iron into a test tube. Heat until the excess sulfur is burned away and the contents begin to glow. Allow the test tube to cool somewhat; then break the test tube by immersing the hot end into cold water. Now repeat Exp. 1 with some of the contents. Results?

Copy the following in your notebook, supplying the word or words needed to complete each statement. Underline the words you supply.

Observations and Questions on Experiments

1. The solubility of sulfur in carbon disulfide is used as an identifying property. Two indications of the dissolving of sulfur are Sulfur dissolves in carbon disulfide.

2. Sulfur (is, is not) —?— attracted to a magnet.

3. Iron (is, is not) —?— attracted to a magnet.

4. The magnet is used as an identifying property of iron.

5. When a magnet is placed near a mixture of iron and sulfur, the magnet attracts the iron.

6. When the mixture of iron and sulfur is shaken with carbon disulfide, the carbon disulfide dissolves the sulfur.

7. When iron and sulfur are mixed, the ingredients of the mixture (lose, retain) —?— their original properties.

8. When the mixture of iron and sulfur is heated, we see evidences of a (physical, chemical) —?— change. Two indications of this kind of change are ~~changes~~ *changes*.

9. After heating the mixture of iron and sulfur, the identifying properties of the iron and sulfur are (lost, retained) —?—.

10. When iron and sulfur are heated together, (an element, a compound, a mixture) —?— called —?— is formed.

Conclusions

1. All substances may be classified according to composition as —?—.

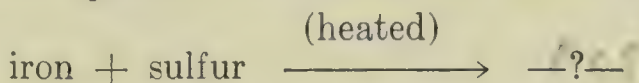
2. In a mixture the ingredients —?— their original properties.

3. In a compound the ingredients —?— their original properties.

4. A mixture is the result of a —?— change.

5. A compound is the result of a ~~change~~ *change*.

6. The word equation for the reaction of iron and sulfur is:



Supplementary Exercises

1. An element is a substance which cannot be —?— by the ordinary types of chemical change into —?—. The two classes of elements are —?—.

2. The oxygen in air retains its original properties. Air is therefore a —?—.

3. The oxygen of water has lost its original properties. Water is therefore a —?—.

4. In the formation of a compound, the ingredient elements are always united in definite proportions by weight. This is a statement of the law of —?—.

5. A mixture (does, does not) —?— conform to the law referred to in Ex. 4 above.

6. Copy the following and after each write the word *element*, *compound*, or *mixture* to indicate its nature: iron, mercuric oxide, oxygen, milk, carbon dioxide, soil, zinc, sugar, salt water, copper sulfide.

7. Copy both List A and List B in your notebook. Write alongside the items of List A the number of the item in List B which corresponds. Do not write any number more than once.

*List A**List B*

- | | |
|-----------------------------------|--|
| (a) Air | 1. A form of energy |
| 3 (b) Gold | 2. Looked for the "philosopher's stone" and the "elixir of life" |
| 6 (c) Gas, liquid, solid | 3. Element |
| (d) Heat | 4. Ninety-two of them |
| 5 (e) Carbon dioxide | 5. Compound |
| 7 (f) Law of Definite Proportions | 6. Physical states of matter |
| 1 (g) Matter | 7. The composition of a compound never varies |
| 8 (h) Sulfur can burn in air | 8. Mixture |
| 2 (i) Alchemists | 9. Anything which occupies space and possesses weight or mass |
| 4 (j) Elements | 10. Chemical property |
8. The four "elements" of the ancients were: fire?—, —?—, —?—, —?—.

9. When two or more elements are chemically united, the substance formed is said to be a —?—.

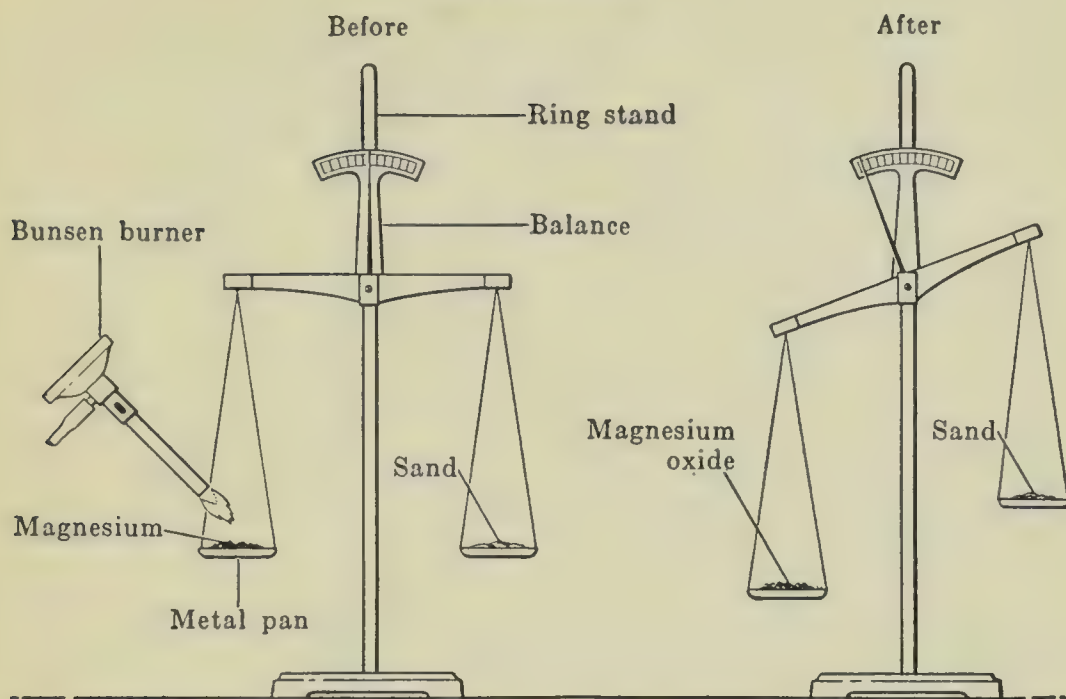
10. The Law of Conservation of Matter states that: —?—.

Optional Questions

1. Name 10 materials which are used in the construction of a house. Classify these as elements, compounds, and mixtures.
2. Define the term "matter." What are the three states of matter?
3. What are the distinguishing properties of each of the three states of matter?
4. Define the term "energy." Name five different forms of energy.
5. Explain the differences between a scientific law, a theory, and an hypothesis. Cite illustrations.
6. How many elements are believed to exist? Which is the most abundant element? Which is the lightest? Which is the heaviest? Which is the most abundant metal?

UNIT 3. EXPLANATION OF BURNING

Experiments



Apparatus for Determining an Explanation of Burning

NOTE: *It is suggested that Experiment 1 be demonstrated.*

1. Use a beam balance or replace one of the pans and the supporting strings of a horn pan balance with an improvised larger metal pan (sheet iron or copper) supported by wire strings as in the figure above. Line this metal pan with asbestos. Place on it a good-sized mound of powdered magnesium. Counterbalance with fine sand. Ignite the magnesium very carefully and completely with a Bunsen flame *from above*. (Avoid losing too much of the product in the form of smoke.) Allow to cool. Result?

NOTE: *If the magnesium is mixed with about one-third its bulk of iron filings, loss due to escape of smoke is minimized. A wad of steel wool may be used instead of magnesium.*

Alternate procedure. Fill a clean porcelain crucible half full with powdered magnesium. Place the crucible, its contents, and a crucible cover on one pan of a balance and counterbalance

with fine sand. Then place the crucible on a pipestem triangle placed on a ring stand. Heat *in air* (burn) thoroughly for about 10 minutes. With the aid of a pair of tongs, use the crucible cover to avoid the loss of the product as smoke. (Do not cover for more than an instant.) When the crucible is cool, replace the crucible and the cover on the balance. Result?

2. Place enough mercuric oxide (the red powder produced when mercury is burned or heated in air) to fill a hard glass test tube to a depth of $\frac{1}{4}$ inch. Hold with a test tube holder in an almost horizontal position and heat, gently at first, and then strongly for a few minutes. While heating, insert a glowing wood splint into the mouth of the test tube. Result? Remove the test tube from the flame and with a wood splint gather together some of the deposit collected on the cool sides of the test tube. What does this deposit resemble?

Copy the following in your notebook, supplying the word or words needed to complete each statement. Underline the words you supply.

Observations and Questions on Experiments

1. When magnesium (or iron) is burned, there is (a gain, a loss, no change) —?— in weight. This shows that magnesium (or iron) —?~~is~~—.

2. When mercury is burned (heated in air), a red powder is produced. This red powder weighs (more than, less than, the same as) —?— the original mercury.

3. When this red powder is decomposed by heating, the gas evolved is —?—. This gas causes a glowing splint to —?—.

4. The substance other than a gas formed on heating the red powder is ~~H₂~~—.

5. The name of the red powder is ~~H₂O~~—.

6. The oxygen from this red powder originally came from —?—.

Conclusions

1. When magnesium or mercury or other similar substances

are burned, or heated in air, they —?— weight because they combine with —?— from the —?—.

2. Ordinary burning is the combination of a substance with —?— accompanied by noticeable heat and light.

3. When an element (metal or nonmetal) is burned in air, the compound formed is called an —?—.

4. Copy and complete the following word equations:

magnesium + oxygen \rightarrow *magnesium oxide*

mercury + oxygen \rightarrow *mercuric oxide*

mercuric oxide \rightarrow

Supplementary Exercises

1. The chemist who first showed that substances —?— weight when burned and who first explained the nature of burning (1774) was —?—. His classic experiment was performed with the metal —?—.

2. It is difficult to prove that the products of burning (or combustion) of coal, wood, or paper weigh more than the original substances because —?—.

3. An ancient erroneous theory of burning which said that a substance escaped from burning things was called the —?— theory.

4. An electric light bulb is either evacuated or filled with an inert gas like nitrogen or argon to —?—.

5. When a substance is decomposed into simpler substances, the process is called —?—.

6. When two or more substances are combined to form one substance, the process is called —?—.

7. Neither sand nor water can burn because —?—.

8. The lowest temperature at which a substance will begin to burn is called its —?—.

9. Most of the common fuels contain carbon and hydrogen either free or combined. When these are burned, the two products formed are —?—.

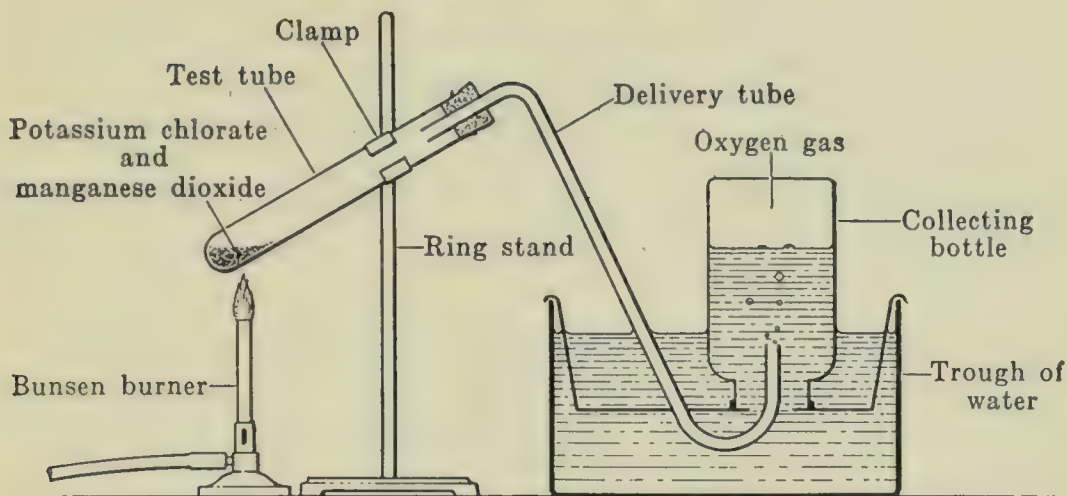
10. A substance can burn if it can —?—. Iron (can, cannot) —?— burn because it —?—.

Optional Questions

1. Briefly describe Lavoisier's experiment which resulted in the first true explanation of burning.
2. What is "flame"? What kinds of substances produce flames when burning?
3. Electric light bulbs have their air removed. Why? Why are some evacuated bulbs filled with argon gas? Why are argon-filled bulbs preferable to evacuated bulbs?
4. Blowing on a candle flame puts it out. Blowing on the embers of a camp fire makes it burn more brightly. Explain.
5. Why do you strike a match in order to ignite it?
6. What are the conditions necessary in order for burning to take place?
7. Why will iron filings ignite in a Bunsen flame while a large piece of iron under the same conditions will not?
8. Explain briefly the principles involved in extinguishing fires.
9. Explain what is meant by the scientific method of solving problems. What are the various steps in the scientific method? Can this method be applied to problems other than those of science? Explain.

UNIT 4. OXYGEN

Experiments



Apparatus for Preparing Oxygen

1. Arrange apparatus as shown in the figure above. On a sheet of paper mix 3 parts of potassium chlorate (12 g.) and 1 part of manganese dioxide (6 g.) sufficient to fill a test tube $\frac{1}{2}$ full. Place the mixture in the test tube and arrange to collect 4 bottles of the gas. *Heat gently.* Do not collect the first few bubbles of gas.

NOTE: When you stop heating the test tube, remove the delivery tube from the water in the trough.

2. Insert a glowing wood splint into the first bottle of oxygen. Result?

3. Ignite a small piece of charcoal (carbon) holding it with a pair of forceps or a piece of wire. Insert the glowing charcoal into the second bottle of oxygen. Result? Remove the charcoal. Add a little limewater and shake. Result?

4. In a combustion spoon, ignite a small piece of sulfur. Then lower the burning sulfur into the third bottle of oxygen. Result?

5. Using the fourth bottle of oxygen, repeat Exp. 4 using red phosphorus instead of sulfur. Result?

6. *Demonstration.* Fill a test tube about $\frac{1}{4}$ full with potas-

sium chlorate crystals (8 g.). Support the test tube in a vertical position on a ringstand. Heat gently until the solid melts and bubbles begin to rise in the liquid. Then heat rather strongly. Test the escaping gas at the mouth of the test tube with a glowing wood splint. Then, **standing at arm's length**, drop an unlit wood splint into the melted potassium chlorate. Results?

Copy the following in your notebook, supplying the word or words needed to make each statement complete. Underline the words you supply.

Observations and Questions on Experiments

1. In preparing and collecting oxygen in this experiment, the first few bubbles of gas should not be collected because —?—.

2. The delivery tube is removed from the water in the trough when the heating is stopped to avoid —?—.

3. This method of collecting oxygen is called the —?— method. The physical property that makes possible this method of collection is the —?— of oxygen in $\frac{14}{16}$?

4. Substances burn more vigorously in oxygen than in air because —?—.

5. When an element is burned in oxygen or in air, an —?— of the element is formed.

6. To test for carbon dioxide, —?— is added and the result is —?—.

7. Oxygen can be prepared by heating —?— alone.

8. The manganese dioxide in this experiment has the effect of —?—. It can be shown that the manganese dioxide (changes, does not change) —?— during the reaction.

9. The wood splint —?— in the presence of melted potassium chlorate because —?— is given off. The potassium chlorate acts as an —?— agent.

10. Copy and complete the word equations:

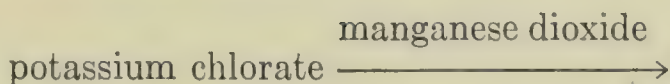
carbon + oxygen \rightarrow

sulfur + oxygen \rightarrow

phosphorus + oxygen \rightarrow

Conclusions

1. The word equation for the laboratory preparation of oxygen is:



2. Four physical properties of oxygen are: colorless, odorless, tasteless, and slightly heavier than air.

3. The outstanding chemical property of oxygen is its ability to combine readily with most elements to form oxides of these elements.

4. Manganese dioxide in this experiment is a catalyst. A catalyst is a substance which speeds up a reaction without being changed.

5. An oxidizing agent, such as potassium chlorate, is a substance which gives off oxygen.

Supplementary Exercises

1. It has been shown that in order for burning to occur a supply of oxygen must be available. Two other essentials for burning are fuel and heat.

2. Carbon dioxide, sulfur dioxide, hydrogen oxide (water), asbestos, lime, etc., will not burn because they do not combine with oxygen.

3. Iron and other metals are used in fire proof construction not because they are unburnable but because they are covered with a non-combustible material.

4. Oxygen for commercial use is obtained from air.

5. Another commercial source of oxygen is water.

6. Oxygen was discovered in 1774 by Lavoisier.

7. Three commercial uses for oxygen are: respiration, welding, and steel making.

8. Rusting of iron is mainly a combination of iron with oxygen from the air. Four substances used to coat iron to prevent rusting are: oil, paint, zinc, and chrome.

9. Fires can be extinguished by removing oxygen from the burning substance, or by lowering temperature below ignition point, or by removing fuel.

10. An allotropic form of oxygen is ozone.

11. The oxygen that is part of water is (combined, uncombined) combined; the oxygen in air is uncombined.

12. The lowest temperature at which a substance will begin to burn is called its ignition point.

13. Spontaneous combustion is an active burning caused by the —?— heat of a —?—.

14. Copy and complete the word equations:

calcium + oxygen \rightarrow CaO

hydrogen + oxygen \rightarrow H_2O

iron + oxygen \rightarrow

15. Copy and complete the following word equations to illustrate three ways of preparing oxygen.

mercuric oxide \rightarrow Hg

water $\xrightarrow{\text{electricity}}$

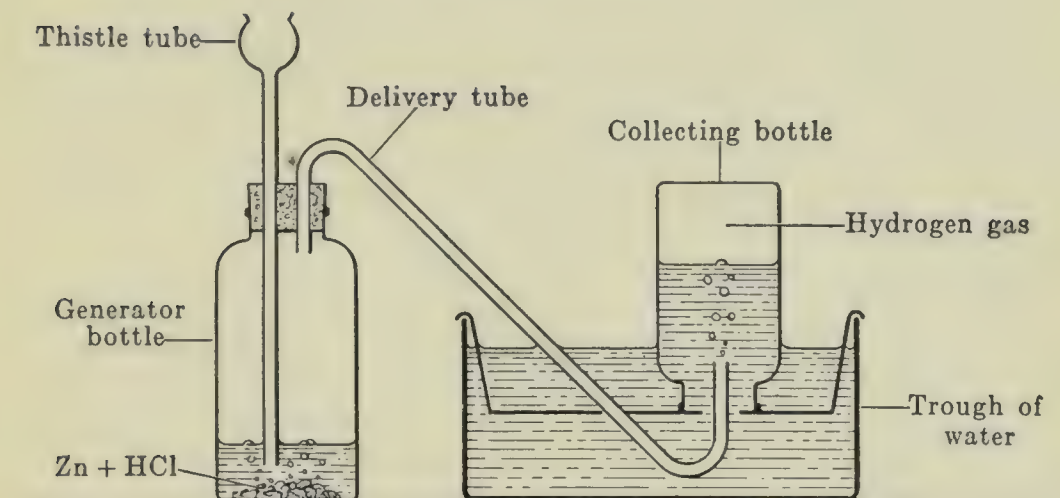
sodium peroxide + water \rightarrow

Optional Questions

1. How does a fish get the oxygen needed for breathing?
2. Explain why exercising will make you "get out of breath."
3. Are all oxides oxidizing agents? Are all oxidizing agents oxides? Explain.
4. Distinguish between burning and combustion.
5. Briefly discuss an industrial chemical process which is dependent upon the aid of a catalyst.
6. What is the difference between a catalyst and an enzyme?
7. Cite some examples of spontaneous combustion. What are the causes? Why should paint-covered rags be kept in metal containers?
8. List five processes, which occur about us every day, that are examples of oxidation.
9. Why will iron rust more rapidly in a moist climate than in a dry climate?
10. When pieces of carbon exactly the same size are burned in air and in oxygen, are the products the same? Is the same amount of heat produced in each case? Is the same temperature produced in each case? Explain.
11. How is ozone prepared? How does it differ from oxygen? What are its uses?
12. Dividing a substance into small particles (subdividing) changes its kindling temperature. In what way? Why?

UNIT 5. HYDROGEN

Experiments



Apparatus for Preparing Hydrogen

Caution: Keep flames away from the hydrogen generator.

1. Arrange apparatus as shown in the figure above. Put into the generating bottle enough mossy zinc (20 g.) to fill $\frac{1}{5}$ of the bottle. Have 4 collecting bottles filled with water inverted in the water trough. Through the thistle tube pour dilute sulfuric acid until the zinc is covered. (If the action in the generator slows down, more acid can be added.) Do not collect the gas given off for at least 2 minutes. Then collect 4 bottles of gas. Cover the bottles with glass plates and place them mouth downward on the table.

NOTE: When you complete Exp. 1, add water to the contents of the generator and empty into a receptacle provided by the instructor.

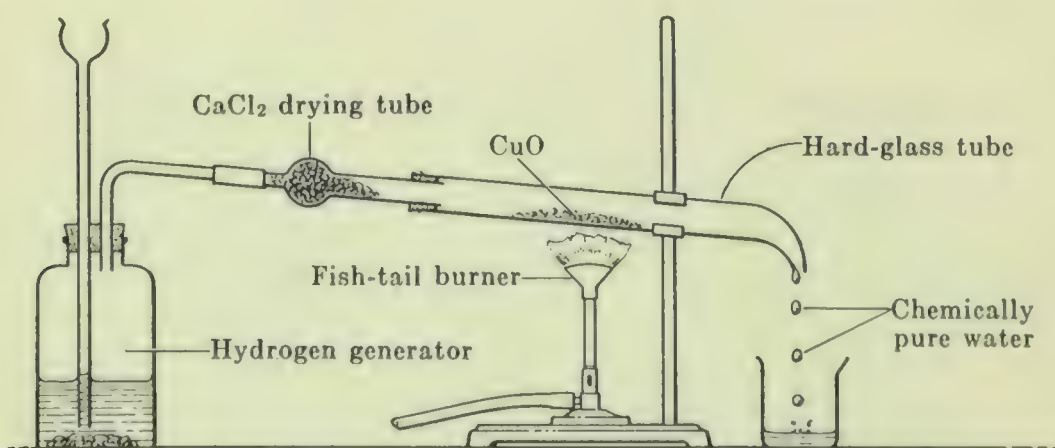
2. (a) Thrust a burning candle into the first bottle of hydrogen, held mouth downward. Result? Remove the candle slowly. Result?

(b) Hold the second bottle of hydrogen uncovered and mouth downward for exactly 1 minute. (Use a watch or clock.) Then thrust a burning splint into the bottle. Result?

(c) Repeat Exp. 2 (b) with the third bottle of hydrogen but hold the bottle mouth upward instead of downward. Result?

(d) Place the fourth bottle of hydrogen, with its glass plate still on, mouth upward on the table. Place a bottle of air on the glass plate just above the bottle of hydrogen. Remove the glass plate. Allow the bottles to stand in this position for exactly one minute. Then bring a burning splint to the mouth of each bottle. Result?

3. *Demonstration.* Filter some of the contents from a hydrogen generator through filter paper. Collect a few drops of filtrate on a watch glass. Evaporate this filtrate over a beaker $\frac{1}{3}$ filled with boiling water. Examine the solid residue on the watch glass. Result?



Apparatus for Reducing Copper Oxide with Hydrogen

4. *Demonstration.* Arrange apparatus as shown in the figure above. Place a little copper oxide in the tube as shown. Do not light the fishtail burner until the hydrogen gas has passed over the copper oxide for at least 2 minutes to drive out the air. Then light the fishtail burner and heat the copper oxide gently until there is a noticeable change in the test tube. Result? (**Keep the flame away from the mouth of the tube.**)

Copy the following in your notebook, supplying the word or words needed to complete each statement. Underline the words you supply.

Observations and Questions on Experiments

1. In preparing and collecting hydrogen the first portion of gas given off should not be collected because —?—.

2. This method of collecting hydrogen is called the —?— method.

3. The hydrogen comes from the —?—.

4. When the burning candle is placed in the bottle of hydrogen, the candle —?—. As the candle is removed from the bottle it is —?— by the burning hydrogen.

5. Hydrogen is (combustible, noncombustible) —?—. It (supports, does not support) —?— combustion.

6. Hydrogen when pure burns (quietly, with explosive violence) —?— to form —?—. When hydrogen is mixed with the element —?— of the air, the mixture burns with explosive violence. The product of this explosion is —?—.

7. Hydrogen is (heavier than, lighter than, as heavy as) —?— air.

8. When the contents of the hydrogen generator are filtered and the filtrate is evaporated on a watch glass, the residue on the watch glass is a —?— colored substance called —?—. (The residue on the filter paper consists of unused zinc and impurities in the zinc, mainly carbon, that were unaffected by the acid.)

9. This residue on the watch glass is not clearly visible in the hydrogen generator because —?—.

10. The calcium chloride drying tube absorbs any —?— mixed in with the hydrogen gas.

11. When hydrogen is passed over hot copper oxide, it combines with the —?— in the copper oxide to form —?—.

Conclusions

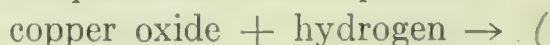
1. The word equation for the laboratory preparation of hydrogen is: zinc + sulfuric acid \rightarrow $\text{ZnSO}_4 + \text{H}_2$

2. Five physical properties of hydrogen observed in this experiment are: —?—.

3. The outstanding chemical property of hydrogen is its ability to combine readily with —?— to form —?—.

4. The removal of oxygen from a substance is the chemical change called —?—. A substance such as hydrogen which removes oxygen from other substances by combining with this oxygen is called a —?— agent.

5. Copy and complete the word equation:



Supplementary Exercises

1. Three metals other than zinc that will displace hydrogen from acids are: —?—.

2. Three metals that will not displace hydrogen from acids are: —?—.

3. In the reduction of copper oxide by hydrogen, it can be shown that 1 gram of hydrogen will remove approximately —?— grams of oxygen to form approximately —?— grams of water.

4. —?— is being substituted for hydrogen in dirigibles and other lighter-than-air craft because —?—.

5. Four uses of hydrogen are: —?—.

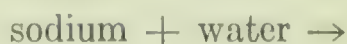
6. —?— is credited with the discovery of hydrogen in 1766 although hydrogen was known by Paracelsus in the sixteenth century.

7. Copy and complete the word equations:



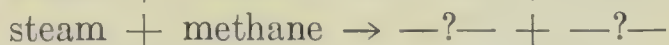
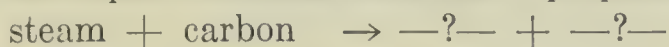
8. All acids contain the element —?—.

9. Copy and complete the following word equations to illustrate other methods of preparing hydrogen:



10. Atomic hydrogen means hydrogen existing as —?—. Ordinary hydrogen is hydrogen existing as —?—. The stable form of hydrogen is —?—. Ordinary hydrogen can be converted into atomic hydrogen by means of —?—. When atomic hydrogen is burned, it produces (more, less) —?— heat than ordinary hydrogen because —?—.

11. Complete the following word equations which represent two recently developed industrial methods of preparing hydrogen.



12. Occlusion is the absorption of gases by —?—.

Optional Questions

1. Compare the physical properties of oxygen and hydrogen. What property makes it possible to collect these gases by the displacement of water method?

2. Describe the hydrogenation of vegetable oils. What catalyst is used? What advantages have the products formed by the hydrogenation of vegetable oils over animal fats?

3. A luminous flame is not produced when pure hydrogen burns. Explain?

4. Explain how the experiment showing the reduction of copper oxide by hydrogen can be used to determine the composition of water by weight.

5. Reduction and oxidation always occur simultaneously. Explain why this must be true.

6. In the laboratory preparation of hydrogen, is it better to use a concentrated or a dilute acid? Why?

7. Why cannot all acids be used to prepare hydrogen?

8. Describe the operation of the oxyhydrogen blow torch. How is atomic hydrogen produced for use in the atomic hydrogen torch? What temperatures can be reached by means of these torches?

9. The elements hydrogen and carbon are present in the combined form in all foods. What happens to these elements in our bodies?

UNIT 6. ACTION OF METALS WITH WATER

Experiments

1. Stand a test tube $\frac{1}{3}$ full of water in a test tube rack or in an empty wide mouth bottle.

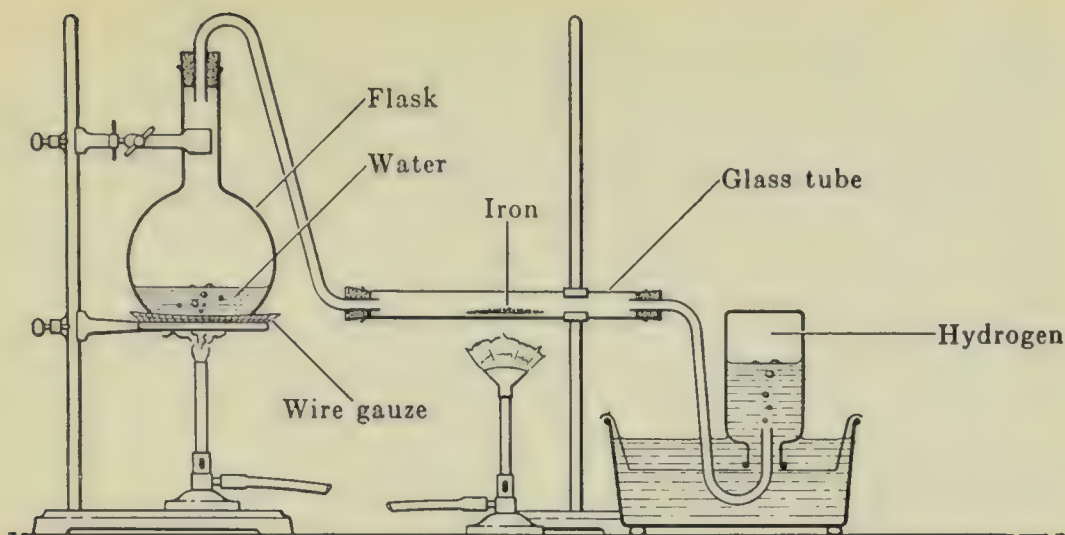
Caution: Stand at arm's length away!

The instructor will now drop a very small piece of sodium (about the size of a small pea) into the test tube. Bring a burning splint to the mouth of the test tube. Result? When the action has stopped, place a drop of the liquid in the test tube on a piece of red litmus paper. Result?

2. Repeat Exp. 1 using a little metallic calcium instead of sodium. If the action does not produce sufficient gas for the test, hold your thumb over the mouth of the test tube to allow some gas to accumulate.

3. Repeat Exp. 1 using a little powdered magnesium instead of sodium. Any action? Carefully heat the test tube until the water in the test tube is almost at its boiling point. Result? Now proceed to collect a test-tubeful of the gas evolved (do not collect the first portion) by displacement of water. **Heat very slowly.** Test the gas with a burning splint. Result? Test the liquid in the test tube with red litmus paper. Result?

4. *Demonstration.* Arrange apparatus as shown at the top of page 21. Incline the Pyrex ignition tube so that the end near the incoming steam is a little lower (to trap any water that may have condensed from the steam). Half fill the flask with water, put some powdered iron in the ignition tube and heat both the iron and the water. Pass steam over the heated iron for a few minutes. Do not collect the first portion of gas liberated. Then collect a bottle full of gas. Test it with a burning splint. Result? Test the product in the ignition tube with red litmus paper. Result? NOTE: If a Pyrex ignition tube is not available, place a steam trap in the setup.



Apparatus for the Reduction of Steam by Iron

Copy the following in your notebook, supplying the word or words needed to complete each statement. Underline the words you supply.

Observations and Questions on Experiments

1. When sodium reacts with water, the gas given off is —?—
This gas comes from —?—.
2. When sodium reacts with water, the product remaining in the water changes litmus from —?— to —?—. This color change shows the presence of a base. The base formed in this experiment is —?—.
3. The action of calcium with water is (more, less) —?— vigorous than the action of sodium with water.
4. Calcium is (more, less) —?— active than sodium.
5. The gas produced when calcium reacts with water is —?—.
6. When calcium reacts with water, the product formed other than the gas, changes litmus from —?— to —?— and is therefore a —?—. The name of this other product is —?—.
7. Magnesium (does, does not) —?— react with water at room temperature.
8. In order to make magnesium react with water, it is necessary to —?— the water.
9. Magnesium is —?— active than calcium.

10. When magnesium reacts with water, —?— and —?— are formed.

11. When steam reacts with iron, the first portion of hydrogen is discarded because it contains —?—.

12. In order to make iron react with water, it is necessary to —?— the iron and use water in the form of —?—. The gas produced in this reaction is —?—.

13. The product, other than hydrogen, formed when iron reacts with steam (does, does not) —?— affect litmus paper. The name of this product is magnetic iron oxide.

14. Iron is —?— active than magnesium.

Conclusions

1. Copy and complete the following word equations:

sodium + water →

magnesium + water →

calcium + water →

iron + water (steam) →

2. A very active metal will react with water at room temperature to produce —?— and —?—.

3. In order to get moderately active metals such as iron to displace hydrogen from water it is necessary to —?—.

Supplementary Questions

1. Cesium, rubidium, and potassium are more active metals than sodium. They would displace —?— from water and react (more, less) —?— vigorously than sodium.

2. In the storeroom metallic sodium is kept under —?— because —?—.

3. Metals can be arranged in the order of their chemical activity. Such an arrangement is called the —?— series.

4. We would not use water to extinguish burning sodium because —?—. We would use —?—.

5. Metals corrode or tarnish in air because they react with the —?—, the —?—, or the —?— of the air.

6. The active and moderately active metals are never found free (uncombined) in nature because —?—.

7. Three metals found in the free state in nature are: —?—.

8. Compounds of the active metals are (stable, unstable) —?— and compounds of the inactive metals are (stable, unstable) —?—.

9. Hydrogen can be released from —?— more readily than from water.

Optional Questions

1. When sodium displaces hydrogen from water, what becomes of the other product formed? How could you obtain it in solid form so that you could examine it?

2. What are bases? What elements do bases always contain? Describe one method of making a base.

3. What is the difference between an oxide and an hydroxide?

4. What coating forms on iron or steel that has been heated and treated with steam? Why is this done commercially?

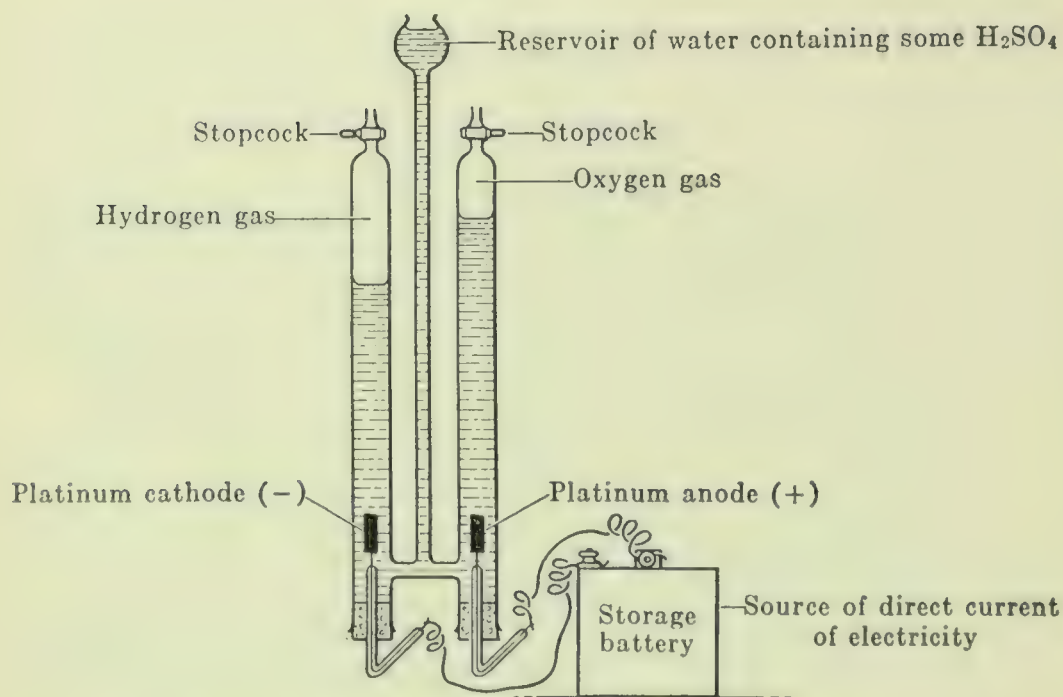
5. What connection is there between the rusting of iron and the presence of moisture in the atmosphere?

6. List 10 metals in their order of activity. Put the most active first.

7. How would you prove that a sample of a colorless liquid was water?

UNIT 7. ELECTROLYSIS OF WATER

Experiment



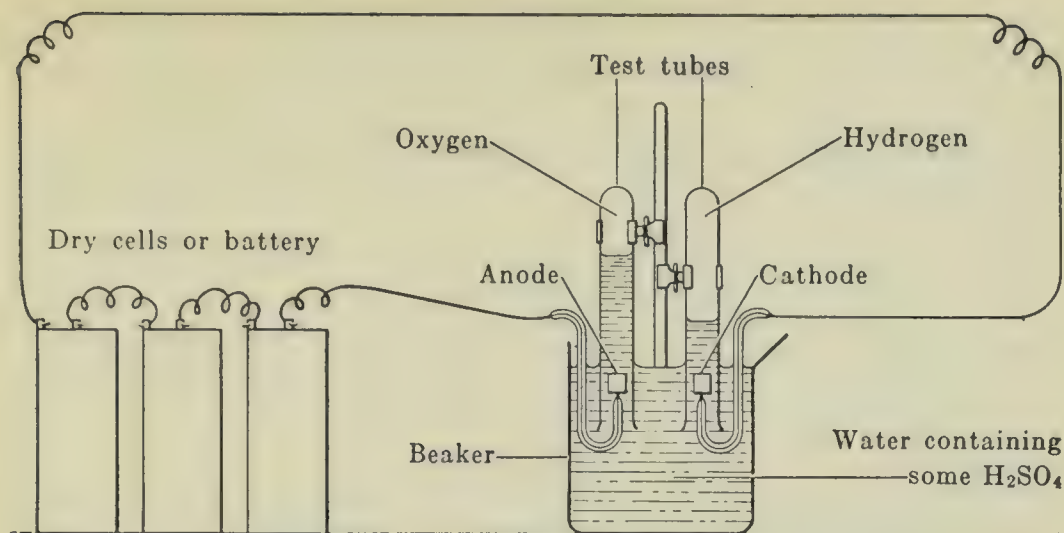
Hoffman Apparatus for the Electrolysis of Water

NOTE: *It is suggested that this experiment be demonstrated by the instructor.*

Arrange either the Hoffman apparatus, shown in the figure above, or the setup shown in the figure on the next page. A 6-volt storage battery, or 3 or 4 dry cells in series, may be used as a source of electricity.

NOTE: *If the Hoffman apparatus is used, the narrow tubes above the stopcocks should be carefully dried by inserting a small piece of filter paper.*

If the other apparatus is used, the electrodes can be made by joining small pieces of platinum wire (1 inch in length) to the copper wires connected to the terminals of the battery. Only the platinum should be in contact with the water. The copper wires can be sealed into glass tubing or coated with paraffine or sealing wax.



Apparatus for the Electrolysis of Water

First fill the apparatus with ordinary water. Connect the battery. Does anything happen? Then add sulfuric acid so that the mixture in the apparatus is approximately 1 part of sulfuric acid to 15 parts of water. (A little sodium fluoride may be used in place of the sulfuric acid.) Connect the battery and pass the current through the mixture. What happens? Continue to pass the current until the tube in which the gas is collecting more rapidly is almost filled with gas. Remove this tube and apply a lighted splint to its mouth. What happens?

Continue to pass the current until the other tube is filled with gas. Remove the tube and apply a glowing splint to its mouth. What happens?

Copy each of the following in your notebook, supplying the word or words needed to complete each statement. Underline the words you supply.

Observations and Questions on the Experiment

1. The function of the sulfuric acid (or sodium fluoride) is —?—. *to conduct the current*

2. Hydrogen is collected at the (positive, negative) —?— electrode. This electrode is called the (anode, cathode) —?—. This electrode is connected to the (positive, negative) —?— pole of the battery.

3. Oxygen is collected at the —?— electrode. This electrode is called the —?—. This electrode is connected to the —?— pole of the battery.

4. Hydrogen is tested with a (burning, glowing) —?— splint. The result is —?—.

5. Oxygen is tested with a —?— splint. The result is —?—.

6. The ratio by volume of the gases produced in this experiment is —?— (volume, volumes) —?— of hydrogen to —?— (volume, volumes) —?— of oxygen.

Conclusions

1. Electrolysis is the decomposition of a substance by means of —?—.

2. In the electrolysis of water, —?— volumes of —?— and —?— volume of —?— are produced.

3. The word equation for the electrolysis of water is:



4. The —?— is liberated at the cathode and the —?— at the anode.

Supplementary Exercises

1. A substance which when placed in water will permit the passage of electricity is called —?—.

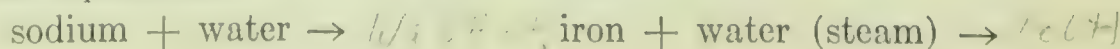
2. The electrodes are made of platinum because —?—.

3. In the formation of water from hydrogen and oxygen, energy is (evolved, absorbed) —?—. In the formation of hydrogen and oxygen from water, energy is (evolved, absorbed) —?—.

4. The —?— is a commercial method of preparing both hydrogen and oxygen.

5. Electricity is a flow of negative particles of electricity called —?—.

6. Other ways than electrolysis of obtaining hydrogen from water are illustrated by the word equations below. Complete these equations.



7. The composition of water by weight is approximately —?— parts by weight of —?— to one part by weight of —?—.

8. "Heavy water," chemically known as —?—, differs from ordinary water in that it contains —?— and therefore has different properties.

Optional Questions

1. Assume air contains 20% of oxygen by volume. 100 cc. of air is mixed with 50 cc. of hydrogen and the mixture is exploded with a spark. What gases combine during the explosion? What volume of each remains after the explosion?

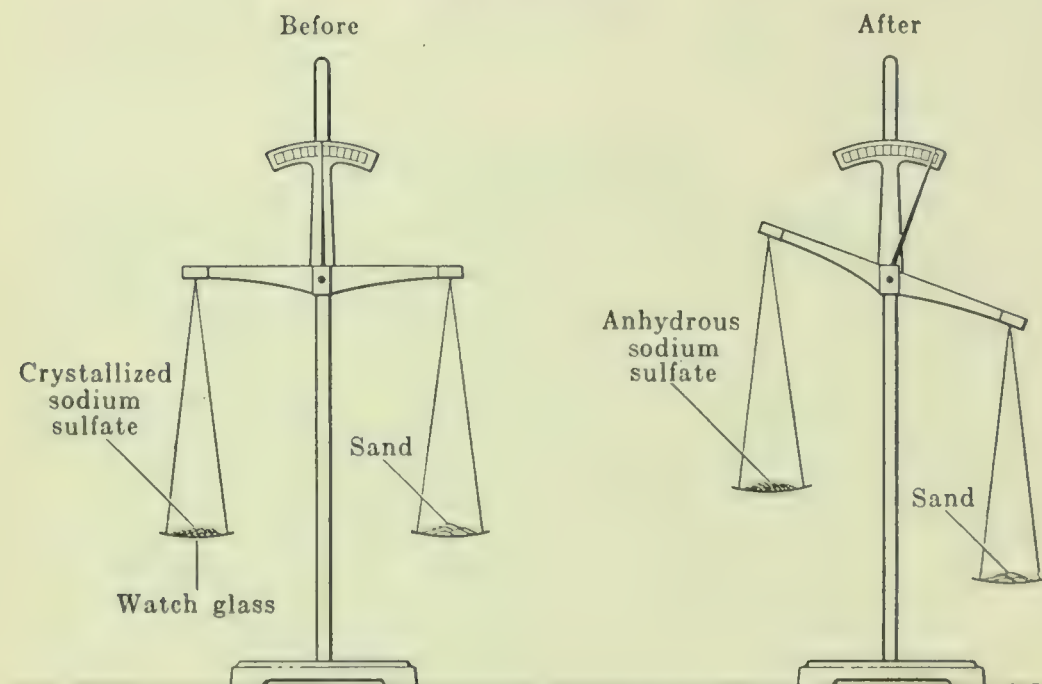
2. Why is the electrolysis of water the commercial method of preparing hydrogen and oxygen in Norway?

3. Explain why the electrolysis of water is called a reversible reaction.

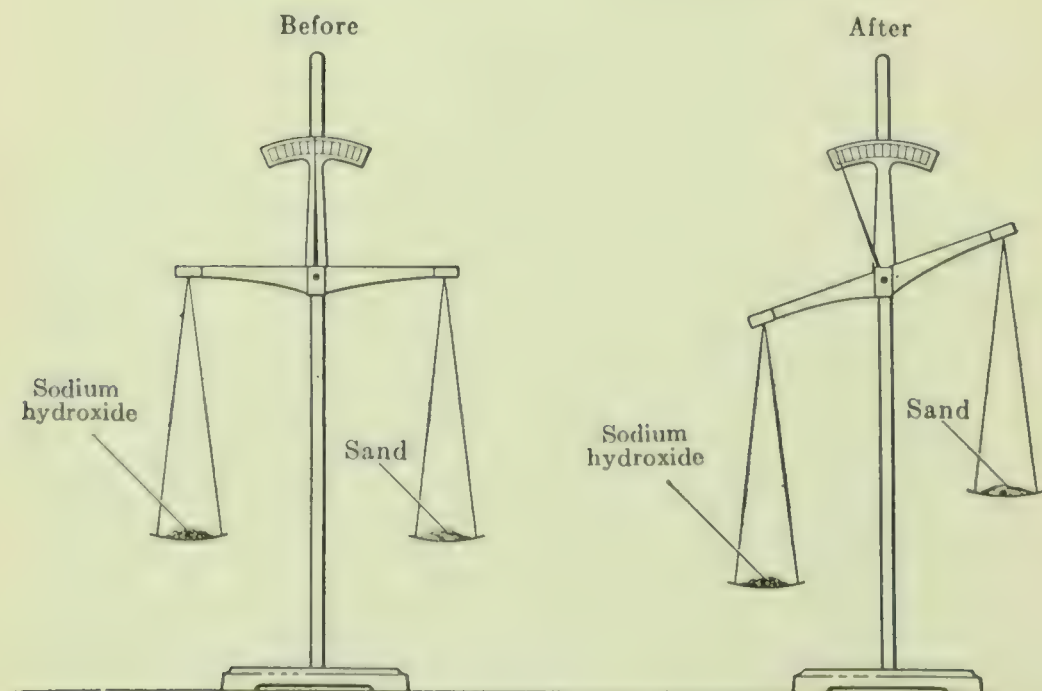
4. Write a report on some of the research work being done today on "heavy water."

UNIT 8. EFFLORESCENCE AND DELIQUESCENT

Experiments



Apparatus for Demonstrating Efflorescence



Apparatus for Demonstrating Deliquescence

1. Place a watch glass containing a mound of sodium sulfate crystals (Glauber's salt) on one pan of a balance and counterbalance with fine sand. On a second balance place a watch glass containing a few pieces of sodium hydroxide and counterbalance with fine sand. Allow both balances to stand undisturbed for at least 30 minutes. Results?

2. In a dry test tube, held in a horizontal position with the mouth a trifle lower than the bottom of the test tube, slowly heat a few small crystals of copper sulfate. Result? Repeat, using in separate dry test tubes a few small crystals of sodium carbonate (washing soda), barium chloride, and potassium nitrate (salt-peter). Is water obtained in each case?

Copy the following in your notebook, supplying the word or words needed to make each statement complete. Underline the words you supply.

Observations and Questions on Experiments

1. The sodium sulfate (increased, decreased) —?— in weight upon exposure to air.

2. The sodium hydroxide (increased, decreased) —?— in weight upon exposure to air.

3. The sodium sulfate crystals changed in appearance to —?—.

4. The sodium hydroxide sticks changed in appearance to —?—.

5. The gain or loss in weight of the above substances when exposed to air is due to the gain or loss of —?—.

6. All crystals (contain, do not contain) —?— water because, as shown by this experiment —?—. The water in certain crystals is known as water of —?—.

7. The following substances, when heated, showed that they were hydrates (contained water of crystallization): —?—.

8. The test tube is held as directed in this experiment because —?—.

9. When the water of crystallization is removed from a crystal the substance changes in appearance to —?—.

Conclusions

1. An efflorescent substance is one which —?— when exposed to air.
2. A deliquescent substance is one which —?—, in sufficient amount to —?—, when exposed to air.
3. An example of an efflorescent substance is —?—.
4. An example of a deliquescent substance is —?—.
5. The water of crystallization can be removed from hydrates by —?—.

Supplementary Exercises

1. The color of anhydrous (without water) copper sulfate is —?—. The color of copper sulfate and water, whether in crystal or solution form, is —?—. This color difference is used as a simple test to show the presence of —?—.
2. Efflorescent and deliquescent substances should be stored in —?—.
3. A dehydrating agent is a substance which —?—.
4. Calcium chloride, being a strongly (efflorescent, deliquescent) —?— substance, is sprinkled on dirt roads and tennis courts in order to —?—.
5. In buying washing soda by the pound it is more economical to buy the (crystals, powder) —?— because —?—.
6. Common table salt cakes in damp weather because of the presence of the impurity —?—, which has strong —?— properties.
7. The water of crystallization in a crystal of a particular substance is present in (a definite, an indefinite) —?— amount.
8. When crystals of calcium sulfate (gypsum) are heated and most of the water of crystallization is driven off, a substance called —?— is formed.
9. Hydration (the taking on of water of crystallization) is responsible for the setting or hardening of —?— and —?—.
10. A hygroscopic substance differs from a deliquescent substance in that —?—. An example of a hygroscopic substance is —?—.

Optional Questions

1. Why are small open containers of calcium chloride often placed inside clocks?
2. Devise an experiment to show that a substance is deliquescent. Hygroscopic. Efflorescent.
3. How could you obtain crystals of copper sulfate from the powdered dry substance?
4. How are crystals formed in nature? What determines the size of crystals as they form?
5. What is meant by the term "amorphous"? Is the normal state of minerals in nature amorphous or crystalline? Explain, citing examples.

UNIT 9. VALENCE, FORMULAS, AND EQUATIONS

TABLE OF IMPORTANT VALENCES

	MONOVALENT	DIVALENT	TRIVALENT
Metals	ammonium NH_4^+ (radical)	calcium Ca^{++}	aluminum Al^{+++}
	hydrogen H^+	copper(ic) Cu^{++}	antimony Sb^{+++}
	mercurous Hg^+	iron (ferrous) Fe^{++}	arsenic As^{+++}
	potassium K^+	lead Pb^{++}	chromium Cr^{+++}
	silver Ag^+	magnesium Mg^{++}	iron (ferric) Fe^{+++}
	sodium Na^+	mercuric Hg^{++}	
		zinc Zn^{++}	
Nonmetals	bromine Br^- (bromide)	oxygen O^{--} (oxide)	nitrogen N^{---} (nitride)
	chlorine Cl^- (chloride)	sulfur S^{--} (sulfide)	phosphorus P^{---} (phosphide)
	fluorine F^- (fluoride)		
	iodine I^- (iodide)		
Radicals	bicarbonate HCO_3^-	carbonate CO_3^{--}	phosphate PO_4^{---}
	chlorate ClO_3^-	sulfate SO_4^{--}	
	hydroxide OH^-	sulfite SO_3^{--}	
	nitrate NO_3^-		
	nitrite NO_2^-		

Copy the following in your notebook, supplying the word or words needed to complete each statement. Underline the words you supply.

I. Complete the following:

1. The metallic elements usually combine with the —?— elements or groups (radicals).

2. The “ous” ending for a metal means a (lower, higher) —?— valence than the “ic” ending.

3. No subscripts are written in the formula when both metal and nonmetal parts have —?— valence.

4. A parenthesis is placed around a radical in a formula if —?—.

5. A chemical radical such as —?— is a combination of —?— which can be considered as behaving like —?— in chemical combination.

6. The number or subscript which is part of a radical (*e. g.* the 4 in SO_4) (changes, does not change) —?— in writing formulas containing this radical.

II. Copy and complete the following table in your notebook indicating the names of the compounds formed by the combination of each metal or metallic radical with each of the nonmetals or nonmetallic radicals listed.

	Sodium	Iron (ferrous)	Iron (ferric)	Ammonium	Hydrogen
Chlorine	Sodium chloride	Ferrous chloride	Ferric chloride	Ammonium chloride	acid
Bromine					acid
Oxygen				×	
Sulfur					acid
Hydroxide					
Carbonate					acid
Bicarbonate			×		acid
Nitrate					acid
Sulfate					acid
Phosphate					acid

III. Write the formulas for each of the following:

- | | |
|------------------------|------------------------|
| 1. Calcium nitrate | 10. Phosphoric acid |
| 2. Ammonium hydroxide | 11. Silver nitrate |
| 3. Ferric oxide | 12. Mercuric oxide |
| 4. Sodium phosphate | 13. Magnesium nitride |
| 5. Zinc chloride | 14. Magnesium nitrate |
| 6. Potassium carbonate | 15. Ferrous sulfate |
| 7. Hydrochloric acid | 16. Potassium chlorate |
| 8. Nitric acid | 17. Aluminum sulfate |
| 9. Sulfuric acid | 18. Hydrogen peroxide |

IV. Copy and complete the table below in your notebook. Write the formulas of the compounds formed by the combination of each metal with each of the nonmetallic radicals listed.

	Chlorine	Bromine	Oxygen	Sulfur	Hydroxide	Carbonate	Bicarbonate	Nitrate	Sulfate	Phosphate
Potassium										
Sodium										
Calcium										
Magnesium										
Aluminum										
Zinc										
Iron (ferrous)										
Iron (ferric)										
Lead										
Hydrogen										
Copper										
Silver					×					
Mercury (ous)					×					
Mercury (ic)						×				
Ammonium (radical)			×							

V. The formulas of the following are correct: SrCl_2 , CaSiO_3 , H_2SO_3 , CoCl_2 , $\text{Na}_2\text{C}_4\text{H}_4\text{O}_6$, Bi_2S_3 , NiCl , RaBr_2 , As_2O_5 , $(\text{NH}_4)_2\text{C}_2\text{O}_4$. Infer the valence of each part and then supply any missing subscripts or parentheses in each of the following. Write the correct formulas in your notebook.

- | | | |
|------------------------|------------------------|---|
| 1. Bi Cl | 3. Na S O ₃ | 5. H C ₄ H ₄ O ₆ |
| 2. H Si O ₃ | 4. Sr N O ₃ | 6. Ra Cl |

- | | | |
|-------------------------------------|--|------------------------|
| 7. Mg Si O ₃ | 10. As Cl | 13. Ni SO ₄ |
| 8. Co N O ₃ | 11. K H C ₄ H ₄ O ₆ | 14. As S |
| 9. Ca C ₂ O ₄ | 12. Bi Si O ₃ | 15. Sr SO ₃ |

VI. How to Write Equations.

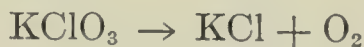
Illustrative example 1. Write the equation for the reaction that takes place when potassium chlorate is heated.

STEP 1. Write the correct formula for the reacting substance.



NOTE: *Most gaseous elements exist alone in diatomic form, that is, 2 atoms to a molecule. Oxygen is written as O₂. The other gases to keep in mind are H₂, Cl₂, and N₂.*

STEP 2. Write an arrow which means yields or produces. Then put down the correct formulas of the products.



STEP 3. Balance the equation.

NOTE: *The number of atoms of each element must be the same on both sides of the equation in order to comply with the law of the conservation of matter.*

(a) Balance one element at a time starting at the extreme left. (A radical may be balanced as a unit if it appears on both sides unchanged.)

(b) If you find an odd number of atoms of an element on one side of the equation and an

Starting with K, we find one atom of K on the left and one atom of K on the right. Similarly we find the Cl atoms are balanced. However we find 3 atoms of O on the left and 2 atoms of O on the right.

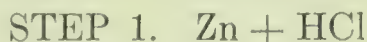
even number of atoms of the same element on the other side, make the odd number even by placing a 2 before the formula which contains the odd number of atoms.

(c) Proceed as before, beginning first with the elements affected by this multiplication by 2.

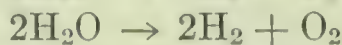
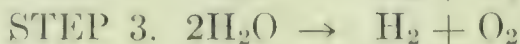
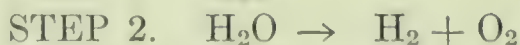
NOTE: (1) *The original correct formulas of the reacting substances must not be changed. Subscripts which are part of the radicals or which are so placed because of valence are not to be changed.*

(2) *All balancing is done by placing numbers in front of the formulas. These numbers are called coefficients.*

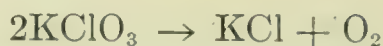
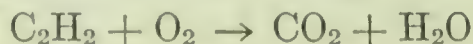
Illustrative example 2. Write the equation for the reaction between zinc and hydrochloric acid.



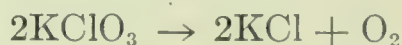
Illustrative example 3. Write the equation for the electrolysis of water.



Illustrative example 4. Write the equation for the complete combustion or burning of acetylene (C_2H_2).

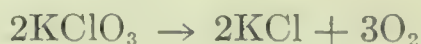


There are now 2 atoms of K on the left and 1 atom of K on the right. Balance as follows:

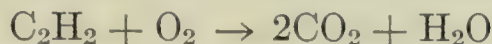


This adjustment makes both the K and the Cl atoms balanced. However, there are 6 atoms of O on the left and 2 atoms of O on the right. Balance this by placing a 3 in front of the O_2 .

The equation is now completely balanced as follows:

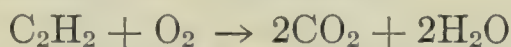


1. Starting with C, we find 2 atoms of C on the left and 1 atom of C on the right. We make the odd number of C atoms even by placing a 2 before the formula CO_2 .

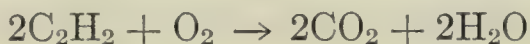


2. There are 2 atoms of H on the left and 2 atoms of H on the right. The H atoms are therefore balanced.

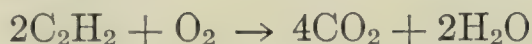
3. The number of O atoms on the left is even (2), but the number of O atoms on the right is odd (5). The odd number of O atoms is in the H_2O ; we therefore make it even by placing a 2 in front of the H_2O as a coefficient.



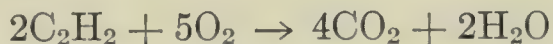
4. The balance of the H atoms has now been destroyed. There are 2 atoms of H on the left and 4 atoms of H on the right. Balance this by placing a 2 in front of the C_2H_2 .



5. This changes the balance of the C atoms. There are 4 C atoms on the left and 2 C atoms on the right. Balance this by changing 2CO_2 to 4CO_2 .



6. The C and H atoms are now balanced. The O atoms are still unbalanced, for there are 2 atoms of O on the left and a total of 10 atoms of O on the right. This can be balanced by placing a 5 in front of the O_2 on the left. The balanced equation now finally reads:



VII. Four Types of Chemical Reactions.

Let us assume:

A is the symbol of a metal.

B is the symbol of a nonmetal.

C is the symbol of another metal.

D is the symbol of another nonmetal.

Let us also assume for the sake of simplicity that the above elements each have a valence of 1.

TYPE OF REACTION	ILLUSTRATION
Direct combination (synthesis)	$A + B \rightarrow AB$
Simple decomposition (analysis)	$AB \rightarrow A + B$
Simple replacement (substitution)	$A + CB \rightarrow AB + C$
Double replacement	$AB + CD \rightarrow AD + CB$

VIII. Write complete, balanced equations (using formulas) for the following reactions. Also indicate the type of reaction illustrated.

- | | |
|----------------------------------|-------------------------------|
| 1. Heating of mercuric oxide | 13. Sodium peroxide + water |
| 2. Iron + sulfur | 14. Zinc + sulfuric acid |
| 3. Burning of magnesium | 15. Zinc + hydrochloric acid |
| 4. Mercury + oxygen | 16. Magnesium + sulfuric acid |
| 5. Heating of potassium chlorate | 17. Iron + hydrochloric acid |
| 6. Burning of carbon | 18. Sodium + water |
| 7. Burning of sulfur | 19. Iron + water (steam) |
| 8. Burning of phosphorus | 20. Copper oxide + hydrogen |
| 9. Calcium + oxygen | 21. Calcium + water |
| 10. Burning of hydrogen | 22. Magnesium + water |
| 11. Iron + oxygen | 23. Ferric oxide + carbon |
| 12. Electrolysis of water | 24. Zinc oxide + carbon |
| | 25. Ferric oxide + aluminum |

UNIT 10. QUANTITATIVE CHEMISTRY

(See also Experiment 4, Unit 18.)

Experiments

EXPERIMENT A. WATER OF CRYSTALLIZATION

(Quantitative Determination)

NOTE: *This experiment should not be performed until the class has learned to write simple formulas and has learned to compute molecular weights from formulas.*

OBJECT:

I. To find the percent of water in crystallized barium chloride.

II. To find the number of molecules of water per molecule of barium chloride.

PROCEDURE: Copy the table below and the one on page 40 in your notebook. Record all weights and indicate all computations in your tables.

1. Carefully weigh a porcelain crucible.

2. Add about 3 grams of pure crystallized barium chloride and weigh the crucible and its contents carefully.

3. Place the crucible on a pipe stem triangle supported on a ring stand just above the hottest part of the Bunsen flame.

4. Heat gently at first. Then heat strongly for at least 15 minutes. Allow to cool and then weigh the crucible and its contents.

5. Heat to constant weight. This means heating for a few minutes, cooling and weighing and repeating this procedure until 2 successive weighings are the same. Record this constant weight.

DATA

A	Weight of empty crucible	g.
B	Weight of crucible + crystallized barium chloride	g.
C	Constant weight of crucible + barium chloride (after heating)	g.

CALCULATIONS

D	Weight of crystallized barium chloride (B - A)	g.
E	Weight of anhydrous barium chloride (C - A)	g.
F	Weight of water driven off (water of crystallization (D - E))	g.

Copy and complete the following computations:

$$\begin{aligned} \text{Percent of water of crystallization} &= \frac{\text{weight of water}}{\text{weight of crystallized BaCl}_2} \times 100 \\ &= \frac{F}{D} \times 100 = \text{---}\% \end{aligned}$$

Number of molecules of water of crystallization = x

Assume formula of crystallized salt = $\text{BaCl}_2 \cdot x\text{H}_2\text{O}$

$$\begin{aligned} \frac{\text{Molecular weight of anhydrous BaCl}_2}{\text{Weight of anhydrous BaCl}_2} &= \\ \frac{x \times \text{molecular weight of water}}{\text{Weight of water}} &= \end{aligned}$$

$$\frac{208}{E} = \frac{18x}{F}$$

$$x = \frac{208 \times F}{18 \times E} = \text{---}\text{---}$$

Correcting to the nearest whole number $x = \text{---}\text{---}$

Copy the following in your notebook, supplying the word or words needed to make each a complete statement. Underline the words you supply.

Conclusions

1. The percent of water of crystallization in crystallized barium chloride is $\text{---}\%$.

2. The number of molecules of water of crystallization (corrected to the nearest whole number) is --- .

3. The formula of crystallized barium chloride is --- .

Observations and Questions on the Experiment.

1. The explosive escape of the water of crystallization may cause —?—.
2. The crucible and its contents should not be weighed while hot because —?—.
3. Crystallized barium chloride was chosen for this experiment because —?—.
4. The computed number of molecules of water (x) is corrected to the nearest whole number because —?—.

EXPERIMENT B. WEIGHT OF 22.4 LITERS OF OXYGEN

OBJECT:

I. To determine the weight of 1 liter of oxygen at standard conditions of temperature and pressure.

II. To determine the weight of 22.4 liters of oxygen at standard conditions (gram-molecular-weight).

PROCEDURE: Copy the tables on page 43 in your notebook. Record all observations in your tables.

1. On a sheet of paper, mix thoroughly about 12 grams of potassium chlorate and 6 grams of manganese dioxide. Place this mixture in a pyrex test tube. Weigh the test tube and contents to .01 gram. Record in the tables in your notebook. Set up apparatus as shown in the figure on page 11, using a clean $2\frac{1}{2}$ liter acid bottle for the collecting bottle.

2. Heat the test tube gently and collect oxygen to fill about three-fourths of the bottle. Do not heat too strongly. Why?

3. Remove the delivery tube. Record the temperature of the water. (This is the same as the temperature of the oxygen.)

4. Raise or lower the bottle so that the water levels inside and outside are the same. (Be careful not to raise the mouth above the surface of the water.) Why? Now place a glass plate over the mouth of the bottle. Remove the bottle (containing water and oxygen) and set it mouth upward on the table.

5. By means of a graduated cylinder determine the volume of water needed to fill the bottle completely. This volume is equal to the volume of the displaced oxygen in the bottle. Record results in the table.

6. When the generating test tube has cooled, remove the stopper and delivery tube and weigh the test tube and its contents again. Record results in the table.

7. Record the barometric pressure. Record the aqueous tension.

DATA

A	Weight of test tube and contents before heating	g.
B	Weight of test tube and contents after heating	g.
C	Volume of oxygen evolved	cc.
D	Temperature of oxygen	°C.
E	Barometric pressure	mm.
F	Aqueous tension	mm.

CALCULATIONS

G	Weight of oxygen evolved (A—B)	g.
H	Volume of oxygen (C)	cc.
I	Pressure of oxygen (E—F)	mm.
J	Temperature of oxygen (D)	°C.
K	Volume of oxygen corrected for standard conditions (Apply Boyle's Law and Charles' Law)	cc.
L	Volume of oxygen (at standard cond.) expressed in liters	l.
M	Weight of 1 liter of oxygen at standard conditions $\left(\frac{G}{L}\right)$	g.
N	Weight of 22.4 liters of oxygen at stand. cond. (M×22.4)	g.
O	Molecular weight of oxygen (textbook) (Weight of 22.4 liters at standard conditions)	g.

Copy the following in your notebook, supplying the word or words needed to make each a complete statement. Underline the words you supply.

Conclusions

1. The weight of 1 liter of oxygen at standard conditions was found to be —?— grams. The true weight of 1 liter of oxygen (textbook) is equal to —?— grams.

2. The weight of 22.4 liters of oxygen at standard conditions was found to be —?— grams. The true weight (textbook) is equal to —?— grams.

3. The weight of 22.4 liters of oxygen at standard conditions is equal to the —?— of oxygen expressed in grams.

Observations and Questions on the Experiment

1. Complete the equation: $\text{KClO}_3 \xrightarrow{\text{MnO}_2}$

2. We must not heat the generating test tube too strongly because we want to avoid —?—.

3. The water levels inside and outside of the bottle containing the oxygen are adjusted so that the pressure of the oxygen inside the bottle is —?— the —?— outside the bottle.

4. To determine the volume of oxygen at standard conditions (—?—° C. temperature and —?— mm. pressure) we must correct the volume obtained in accordance with —?— Law and —?— Law.

EXPERIMENT C. EQUIVALENT WEIGHT OF MAGNESIUM

OBJECT: To determine the weight of magnesium that displaces one gram of hydrogen.

PROCEDURE: Copy the table on page 46 in your notebook. Record all observations in your tables.

1. Start with a battery jar about $\frac{2}{3}$ full of water *at room temperature*.

2. Weigh a piece of clean magnesium ribbon (between .04 and .045 grams) accurately. Record the weight in the table in your notebook. Enough pieces for a class can be obtained by weighing a uniform strip of magnesium ribbon a few meters long and cutting off proportionate lengths.

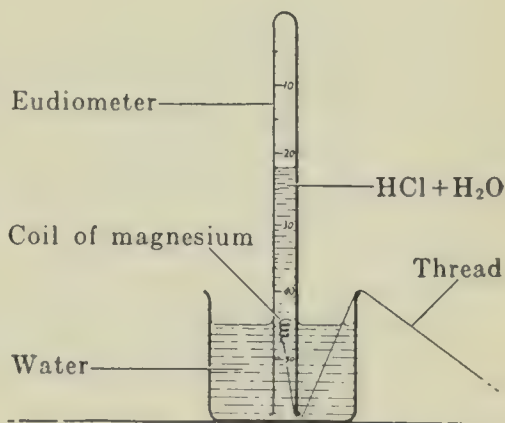
3. Roll the piece of magnesium into a small loose coil, so that it will fit into the eudiometer (gas measuring) tube. Tie a thread to the coil.

4. Pour about 5 cc. of concentrated hydrochloric acid into the eudiometer tube. Fill the tube with water very carefully to avoid undue mixing.

5. Lower the magnesium coil about 2 inches into the eudiometer tube holding on to the thread so that the coil will not sink too far.

6. Place your thumb over the mouth of the eudiometer tube and invert the tube in the battery jar of water. Rest the mouth of the tube on the bottom of the battery jar and on the thread as shown in the figure at the right.

7. The acid being heavier than the water will settle to the bottom and react with the magnesium. When the action has stopped, raise or lower the tube so that the level of the liquid inside is at the level of the liquid outside the tube.



8. Record the volume of hydrogen gas obtained. Record the temperature of the water. Record the barometric pressure. Record the aqueous tension.

DATA

A	Weight of magnesium ribbon used	g.
B	Volume of hydrogen obtained	cc.
C	Temperature of the water	°C.
D	Barometric pressure	mm.
E	Aqueous tension	mm.

CALCULATIONS

F	Volume of hydrogen (B)	cc.
G	Pressure of hydrogen (D—E)	mm.
H	Temperature of hydrogen (C)	°C.
J	Volume of hydrogen corrected for standard conditions (Apply Boyle's Law and Charles' Law)	cc.
K	Weight of hydrogen $\left(\frac{J}{1000} \times .09\right)$	g.
L	Weight of magnesium that replaces one gram of hydrogen (equivalent of magnesium) $\left(\frac{A}{K}\right)$	g.

Copy the following in your notebook, supplying the word or words necessary to complete each statement. Underline the words you supply.

Conclusions

1. The equivalent (or combining) weight of an element may be defined as the number of grams of that element that will combine with or replace —?—.

2. The equivalent (or combining) weight of magnesium was found to be —?—.

3. The true equivalent (or combining) weight of magnesium (textbook) is equal to —?—.

Observations and Questions on the Experiment

1. Complete the equation: $\text{Mg} + \text{HCl} \rightarrow$
2. The above is known as a —?— type of reaction.
3. The water levels inside and outside the eudiometer tube are adjusted so that the pressure of the hydrogen inside the tube is —?— the —?— outside of the tube.
4. To determine the volume of hydrogen at standard conditions (—?—° C. temperature and —?— mm. pressure) we must correct the volume obtained in accordance with —?— Law and —?— Law.
5. One liter (l.) is equal to —?— cubic centimeters (cc.).
6. The weight of one liter of hydrogen gas at standard conditions is equal to —?— gram.

UNIT 11. AIR

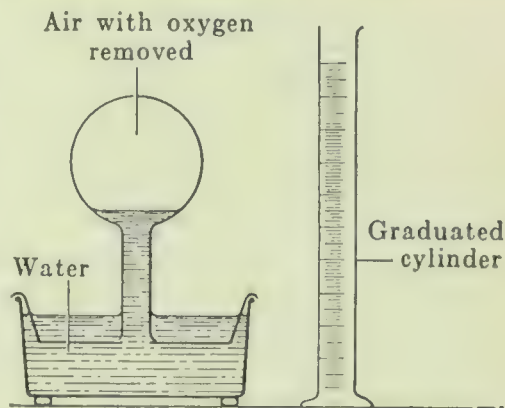
Experiments

NOTE: *It is suggested that these be demonstration experiments.*

1. Percent of oxygen in air (volumetric).

Caution: Yellow phosphorus must be kept and handled under water.

Place a small piece of yellow phosphorus in a dry flask (pyrex glass if available) and seal with a solid rubber stopper. Gently warm the flask with a Bunsen flame until there is evidence of reaction. What happens? When the white smoke has settled and



Apparatus for Determining the Percent of Oxygen in Air

the flask is cool, immerse the mouth of the inverted flask in a battery jar of water as in the figure at the right. Remove the stopper under water. What happens? Adjust the flask so that the level of the water inside is at the level of the water outside. Then while the flask is in this position, replace the stopper securely in the mouth of the flask. Remove the flask. (Remove the stopper and perform Experiment 2 (a). Then measure the volume of water now in the flask by pouring it into a graduated cylinder. Now measure the total volume of the flask by filling it with water (replace the rubber stopper to allow for its volume) and then pour this water into a graduated cylinder. Make a table like the one below in your notebook and enter your results.

Volume of air in flask	cc.
Volume of water that entered flask	cc.
Volume of oxygen removed by phosphorus (Volume of oxygen in the air of flask)	cc.
Percent of oxygen in air (volumetric)	%

Alternate procedure for Exp. 1. In a battery jar of water, float a little red phosphorus either in a small crucible or on a flat piece of cork. Invert a graduated cylinder (1 liter preferred) with two rubber bands stretched around it over the phosphorus so that the edge of the cylinder is just under water. Support the cylinder in a ring attached to a ring stand. Mark the water level on the cylinder by adjusting one of the rubber bands. Raise the cylinder slightly and touch the red phosphorus with a hot wire. Quickly replace the cylinder. Let stand until the fumes subside. (If the water level in the battery jar is about to go below the mouth of the cylinder, add more water.) Mark the new water level with a second rubber band. Complete the table on page 48.

2. Other ingredients of air:

(a) Test the gas remaining in the flask (almost pure nitrogen) by inserting a burning splint. Result?

(b) Place a stick of dry sodium hydroxide on a watch glass exposed to air. Examine at the end of the period and also after 24 hours. Result?

(c) Place some limewater on a watch glass and expose to air. Examine at the end of the period and also after 24 hours. Result?

Copy the following in your notebook, supplying the word or words needed to make each a complete statement. Underline the words you supply.

Observations and Questions on Experiments

1. Before phosphorus is added, the flask contains —?—.
2. The phosphorus combines with the —?— in the flask to form a —?— colored (gas, liquid, solid) —?— named —?—.
3. Phosphorus is used to remove —?— from the confined volume of air because —?—.
4. Copy and complete the equation: $P + O_2 \rightarrow$
5. The oxide of phosphorus produced has (a negligible, an appreciable) —?— volume and is (soluble, insoluble) —?— in water.

6. The water rises in the flask because, when the oxygen is removed, there is created a partial —?—, and the —?— outside forces up the water to fill the —?—.

7. Almost 99% of the gas left in the flask (after the oxygen was removed) is —?—. When tested with a burning splint this gas is shown to be (combustible, noncombustible) —?— and a (supporter, nonsupporter) —?— of combustion.

8. When exposed to air for a number of hours, a stick of sodium hydroxide becomes —?— because it has absorbed —?— from the air. A substance which behaves the way sodium hydroxide does here is called a —?— substance.

9. When limewater is exposed to air, there is formed a —?— because the limewater reacts with the —?— of the air to form the insoluble —?—.

Conclusions

1. In this experiment it is shown that air contains approximately —?— per cent by volume of oxygen.

2. The two gaseous elements —?— make up 99% by volume of air.

3. Two compounds shown to be present in air are —?—.

Supplementary Exercises

1. A test tube with some moist iron filings clinging to its inner surface is inverted with its mouth just below the surface of a dish of water. After some hours the water is observed to have —?— and to have filled —?— of the test tube. The explanation for this is —?—.

2. Dry, pure air contains (by volume) approximately 78% of —?—, 21% of —?—, 0.9% of the inert gas, —?—, and traces of the other inert gases: —?—.

3. The per cent of —?— in air varies greatly.

4. Copy and complete the table below in your notebook.

FOUR IMPORTANT INGREDIENTS OF AIR	A NATURAL PROCESS WHICH REMOVES IT FROM AIR	A NATURAL PROCESS WHICH RETURNS IT TO AIR

5. One good way of proving that air is a mixture and not a chemical compound is to show that it does not conform to the law of —?—. Its composition is not absolutely —?—.

6. The element nitrogen is (heavier, lighter) —?— than oxygen, is relatively (active, inactive) —?—, and its compounds are (stable, unstable) —?—.

7. Nitrogen is an essential constituent of an important class of foods called —?—.

8. Of the gaseous elements found in air two that are diatomic are —?—. Two that are monatomic are —?—.

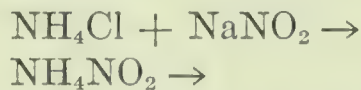
9. Oxygen, nitrogen, neon, and argon are now obtained for industrial use from air in the —?— state. When this —?— air evaporates, the gas that escapes first is mostly —?—.

10. Helium, now obtained largely from —?— found in Texas containing as high as 1% of helium, is used in —?—.

11. In 1892 —?— observed that *pure* nitrogen obtained from various compounds weighed less than the “nitrogen” obtained from air. This hinted that atmospheric “nitrogen” might contain a (heavier, lighter) —?— gas. In 1894, he and —?— removed the nitrogen from the atmospheric “nitrogen” by combining it with heated magnesium to form —?—. There was left an inert gas named —?— which was found to comprise about —?—% of the air.

12. Complete the equation: $\text{Mg} + \text{N}_2 \rightarrow$

13. Pure nitrogen is conveniently prepared by heating a mixture of ammonium chloride and sodium nitrite. Copy and complete the equations:



14. Modern standards of ventilation indicate that the room temperature of air should be about —?— degrees Fahrenheit, that the air (should, should not) —?— be circulating, that a slight excess of carbon dioxide is (harmful, harmless) —?—, that a slight deficiency of oxygen is (harmful, harmless) —?—, and that our comfort to a great extent depends on the relative humidity of the air. Relative humidity is the amount of the —?— as compared to the —?— at the same temperature.

15. The correct relative humidity for comfort is approximately 50%. On a sultry summer day we are uncomfortable mainly because the relative humidity is (too low, too high) —?—.

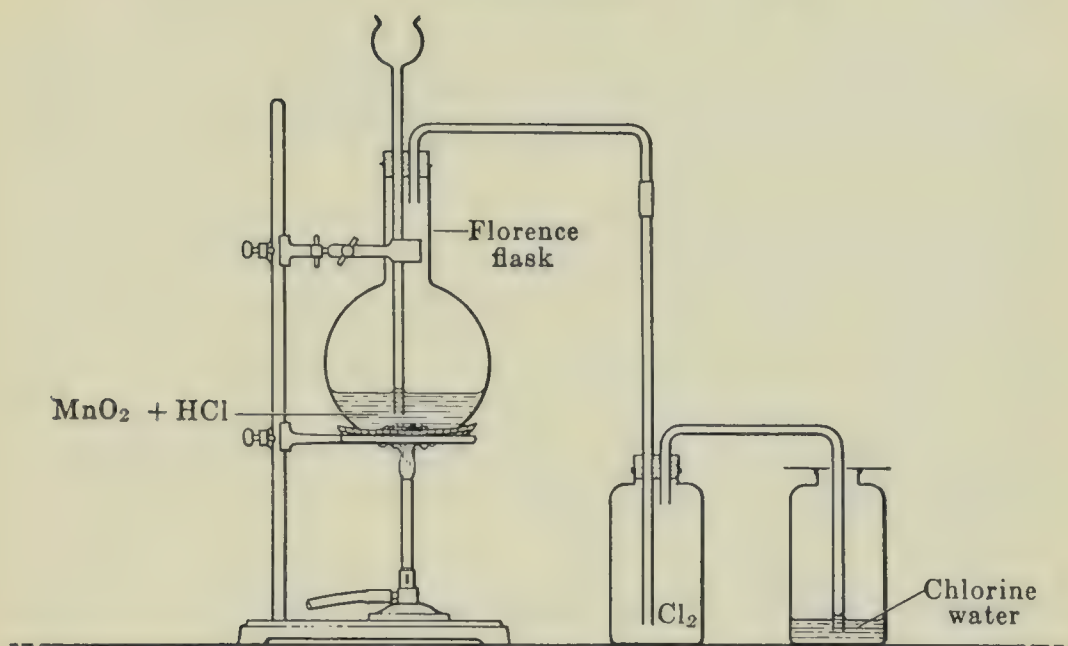
16. In a steam heated room in the winter time we are often uncomfortable because the relative humidity is (too low, too high) —?—. This can be corrected by —?—.

Optional Questions

1. How is air liquefied? Name some uses for liquid air.
2. Explain the oxygen-carbon dioxide life cycle.
3. Explain the nitrogen cycle.
4. What is the original energy source for the above mentioned cycles?
5. Explain the use of argon in electric light bulbs.
6. Explain the use of the barometer in weather prediction.
7. What determines whether the air we breathe is healthful?
8. Explain what is meant by fractional distillation. How is this process applied to air.

UNIT 12. CHLORINE

Experiments



Apparatus for Preparing Chlorine

NOTE: *It is suggested that in view of the poisonous nature of chlorine these experiments be performed by the instructor.*

1. Arrange apparatus as shown in the figure above. (It is suggested that 5 collecting bottles with a saturated solution of sodium thiosulfate, "hypo," in the fifth bottle, be arranged in series.) Place about $\frac{1}{2}$ a test-tubeful of manganese dioxide (24 g.) in the flask. Pour down the thistle tube about a test-tubeful of concentrated hydrochloric acid. Heat very slowly and gently. When the 4 collecting bottles are filled with chlorine gas (how do we know?), remove them and cover with glass plates. Substitute 4 other collecting bottles. Six bottles of chlorine will be needed in these experiments.

2. Pour a test-tubeful of water into the first bottle of chlorine. Place the palm of your hand over the mouth of the bottle and shake vigorously. Any evidence of solution? Place pieces of colored calico, blue and red litmus paper into the bottle. Cover with a glass plate. Result?

3. In the second bottle of chlorine place dry pieces of colored calico cloth, blue and red litmus paper. Compare with action in first bottle.

4. Place a small piece of freshly cut sodium in a combustion spoon and heat for a moment. Drop the sodium into the third bottle of chlorine and cover with a glass plate. Allow to stand. Result?

5. Into the fourth bottle of chlorine sprinkle a little powdered antimony metal. Result?

6. Into the fifth bottle of chlorine insert a burning jet of hydrogen. Result? Blow your breath over the mouth of the bottle. Result?

7. Into the sixth bottle of chlorine insert a burning candle or wax (paraffin) taper. Result? Blow your breath over the mouth of the bottle. Result?

Copy the following in your notebook, supplying the word or words needed to make each a complete statement. Underline the words you supply.

Observations and Questions on Experiments

1. Five bottles in series are often used (4 collecting ones instead of the customary one bottle) because —?—.

2. This method of collecting chlorine is known as the —?— method. Two physical properties of chlorine that make possible this method of collection are: —?—.

3. The purpose of the "hypo" in the last bottle is —?—.

4. We know that the collecting bottles are filled with chlorine gas when —?—.

5. The purpose of the hydrochloric acid in the reaction by which chlorine is prepared is to —?—, and the purpose of the manganese dioxide is to —?—. A substance which serves the purpose of the manganese dioxide in this experiment is called —?—.

6. Chlorine is (soluble, slightly soluble, insoluble) —?— in water as shown by —?—.

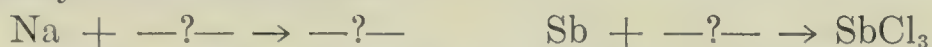
7. The colored piece of cloth was —?— and the litmus paper was —?— by the mixture of chlorine and water.

8. As compared to the effect of chlorine on wet colored cloth and litmus paper, the effect of chlorine on dry colored cloth and litmus paper —?—.

9. In order to bleach with chlorine the presence of —?— is essential.

10. Sodium and antimony, both (metals, nonmetals) —?—, combine very readily with chlorine with the evolution of noticeable —?—.

11. The balanced equations for the reactions of sodium and antimony with chlorine are:



12. Chlorine (supports, does not support) —?— the combustion of hydrogen. The name of the product formed is —?—.

13. The product produced in Ex. 12 above causes —?— when in contact with the moisture of exhaled breath. This shows the product is —?— in water.

14. Chlorine (supports, does not support) —?— the combustion of a paraffin candle or taper. Paraffin is a compound of hydrogen and carbon.

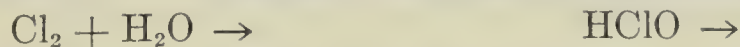
15. The products formed when paraffin is burned in chlorine are —?— and —?— as shown by —?— and by —?—.

16. The experiment with the burning paraffin shows that chlorine has a strong attraction for —?—.

Conclusions

1. The balanced equation for the laboratory method of preparing chlorine is: $\text{MnO}_2 + \text{—?—} \rightarrow$

2. When chlorine reacts with water, the reaction may be written in two stages. Complete the equations:



In bleaching, one of these products, namely the —?— decomposes into hydrochloric acid and —?—. This latter substance actually does the bleaching by combining with the coloring matter to form a —?— substance.

3. Chlorine combines readily with many metals to form salts called —?—.

4. Chlorine combines readily with hydrogen. Write a balanced equation for this reaction.

Supplementary Exercises

1. To counteract some of the effects of inhaling chlorine gas, —?— or —?— is inhaled.

2. Three uses of chlorine are: —?—.

3. Chlorine is made commercially by the —?—. Write the equation for this process.

4. Chlorine water is kept in dark colored bottles because —?—.

5. Two substances other than manganese dioxide that can be used to liberate the chlorine from hydrochloric acid are —?—.

6. Chlorine water and hydrogen peroxide are similar in their bleaching action because —?—.

7. Two textile fibers that should not be bleached with chlorine water are —?— because —?—.

8. Bleaching powder (formula —?—) is made by passing chlorine over —?—. The chlorine is liberated from bleaching powder for bleaching or disinfecting purposes by the addition of —?—.

9. The sodium compound which is similar to bleaching powder has the chemical name of —?—; its formula is —?—. Three commercial names of its water solution, used for —?— and —?— purposes, are —?—.

10. It is more practical to use bleaching powder or liquid chlorine rather than chlorine gas because —?—.

11. Chlorine was discovered by —?— in 1774. It was shown to be an element and named chlorine, however, by —?— about forty years later.

Optional Questions

1. Describe the commercial process for bleaching cotton cloth.

2. Describe the commercial process of preparing chlorine by means of the electrolysis of brine.

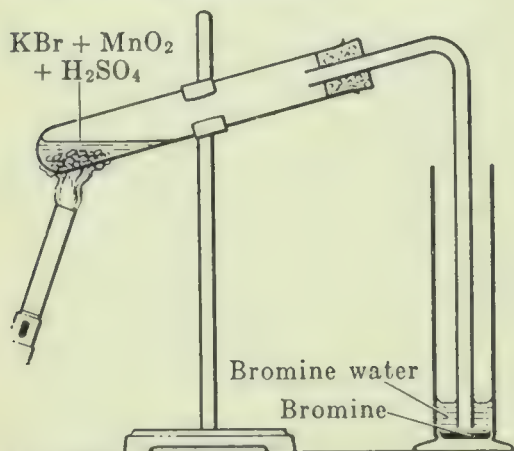
3. Compare the chemical properties of chlorine with those of oxygen.

4. Why must a water solution of chlorine be freshly prepared?
5. What are the halogen elements? Why are they grouped into a chemical family?
6. Why cannot printer's ink be bleached with chlorine?
7. Why is it possible to collect chlorine gas by displacement of brine?
8. Another laboratory method for making chlorine is by the reaction of sodium chloride with sulfuric acid and manganese dioxide. Compare this method with the method used in this unit.
9. Mention three chemical changes that are brought about by means of light energy.
10. Write a report on the importance of chlorine compounds, mentioning at least seven compounds and their uses.

UNIT 13. BROMINE

Experiments

Caution: Bromine is dangerous. Do not get the liquid on your skin or inhale the fumes.



Apparatus for the Preparation of Bromine

1. Mix a small amount (only a few small crystals) of sodium bromide with an equal bulk of manganese dioxide on a sheet of paper. Place the mixture in a dry test tube. Arrange apparatus as shown in the figure at the left. The end of the delivery tube should be just below the surface of the water which half fills the collecting tube. Add 3

cc. of sulfuric acid (2:1) to the mixture in the generating test tube. Heat gently and collect a few drops of liquid bromine in the water.

Remove the delivery tube from the water before you stop heating.

NOTE: The instructor may deem it safer and more desirable to use a glass retort like the one used in preparing nitric acid.

2. Fill a test tube $\frac{1}{3}$ full of water. Add about 4 cc. of carbon disulfide. Which is heavier? Put your thumb over the mouth of the test tube and shake. Are the liquids miscible? To this mixture add a little bromine water from the collecting tube. Shake again and allow to stand. Result?

3. Dissolve a pinch of sodium bromide in a test tube $\frac{1}{6}$ full of water. Add about 4 cc. of carbon disulfide (or carbon tetrachloride). Shake. Allow to stand. Any color imparted to the carbon disulfide? To this mixture add a little chlorine water and shake. Allow to stand. Result? This is a test for any bromide.

4. To $\frac{1}{3}$ of a test-tubeful of bromine water add a small piece of the metal calcium. Any evidence of action?

5. *Demonstration.* Place 2 or 3 drops of liquid bromine in

a beaker. Sprinkle some powdered antimony into the beaker. Result?

Copy the following in your notebook, supplying the word or words needed to make each a complete statement. Underline the words you supply.

Observations and Questions on Experiments

1. The bromine is given off in the ~~—?—~~ state and then condensed to its normal ~~—?—~~ state in cold water. The bromine is collected in the water because ~~—?—~~.

2. Bromine is (insoluble, slightly soluble, very soluble) ~~—?—~~ in water; has a ~~—?—~~ color; is (lighter, heavier) ~~—?—~~ than water, and is a (volatile, nonvolatile) ~~—?—~~ liquid.

3. In the laboratory preparation of bromine the functions of the reacting substances are: sodium bromide: ~~—?—~~; sulfuric acid: ~~—?—~~; manganese dioxide: ~~—?—~~.

4. Complete and balance the equation:



5. Carbon disulfide is (lighter, heavier) ~~—?—~~ than water and is (miscible, nonmiscible) ~~—?—~~ with it.

6. Bromine is (more, less) ~~—?—~~ soluble in carbon disulfide than in water.

7. Bromine when dissolved in carbon disulfide imparts a characteristic ~~red?~~ color to it.

8. Sodium bromide (does, does not) ~~—?—~~ impart a color to the carbon disulfide.

9. When chlorine water is added to the solution of sodium bromide, the free ~~—?—~~ liberated colors the carbon disulfide ~~—?—~~.

10. Chlorine is (more, less) ~~—?—~~ active than bromine.

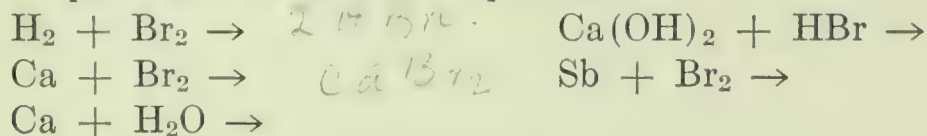
11. When metallic calcium is added to bromine water, the evidence of chemical action is ~~—?—~~.

12. When calcium combines with bromine, the substance formed is ~~—?—~~.

13. When calcium reacts with water, there is produced ~~—?—~~ and ~~—?—~~ gas. This gas would react with bromine to form ~~—?—~~.

14. Antimony combines with bromine with noticeable —?— and —?—. Antimony is a (metal, nonmetal) —?—.

15. Complete and balance the equations:



Conclusions

1. Bromine is prepared in the laboratory by the reaction of —?—. It is collected by —?—.

2. Four physical properties of bromine are: —?—.

3. Bromine (similar to chlorine) combines readily with (metals, nonmetals) —?—. It combines with the —?— of water to liberate —?— which bleaches and disinfects.

4. In testing for a bromide, add —?— and —?— to the solution of the unknown. Shake. A —?— color in the bottom —?— layer indicates that free —?— has been liberated from a bromide.

Supplementary Exercises

1. Carbon disulfide is not colored by the bromine in —?—, but is colored by —?—.

2. Chlorides have a (greater, lesser) —?— heat of formation, are (more, less) —?— stable and will tend to form (more, less) —?— readily than bromides.

3. An important commercial source of bromine is the —?— found in the brine of salt wells and the sea. The bromine is extracted from this compound by replacement with —?—.

4. —?— is the only nonmetallic element that exists in the —?— state at ordinary conditions.

5. Two important uses for bromine are —?—.

6. Silver bromide is used in —?— because it is —?—.

7. —?— bromide is used medically as a sedative.

8. Most of the tear gases used are compounds of —?—.

9. If bromine is accidentally inhaled, the injurious effects can be somewhat lessened by breathing —?—.

10. The halogen elements arranged in their order of activity (most active first) are: —?—.

11. The halogen elements are grouped as a chemical family because —?—.

Optional Questions

1. Compare bromine with chlorine with respect to physical and chemical properties.

2. Describe the commercial process of making bromine.

3. What is a lachrymator? Why mention it in this unit?

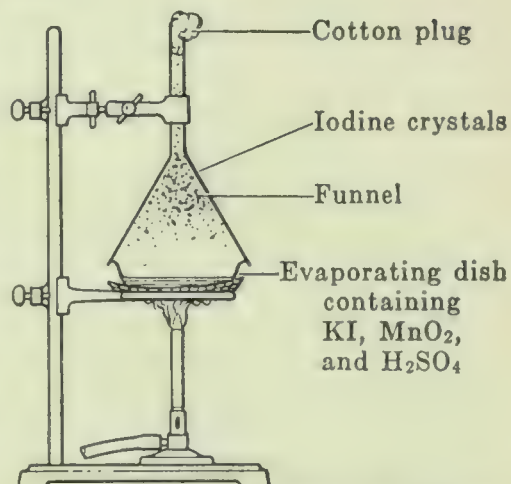
4. Explain how bromine may act as a bleaching agent.

5. Are chlorine and bromine supporters of combustion? What difference, if any, is there between burning and combustion?

UNIT 14. IODINE AND FLUORINE

Experiments

1. Mix a small amount (1 g.) (only a few crystals) of potassium iodide with an equal bulk (2 g.) of manganese dioxide on a sheet of paper. Place the mixture in an evaporating dish. Arrange apparatus as shown in the figure at the right. Add 5 cc. of sulfuric acid (2:1). Heat the dish gently until the iodine formed vaporizes and condenses on the funnel. Remove the funnel and examine the iodine crystals.



Apparatus for Preparing Iodine

2. Fill five test tubes to a depth of one inch with water, alcohol, carbon tetrachloride, carbon disulfide and a water solution of potassium iodide, respectively. Add a little iodine scraped from the funnel with a wood splint to each. Shake each tube. Make a table of solubility like the one below in your notebook and complete it.

TABLE OF SOLUBILITY OF IODINE

SOLUTE	SOLVENT	SOLUBLE, SLIGHTLY SOLUBLE, INSOLUBLE	COLOR OF SOLUTION
Iodine	Water		
Iodine	Alcohol		
Iodine	Carbon tetrachloride		
Iodine	Carbon disulfide		
Iodine	Water solution of potassium iodide		

3. To a cold starch suspension add a drop or two of the solution of iodine in alcohol. Note color. The color change is char-

acteristic and is used as a test for either free iodine or starch.

4. Dissolve a pinch of potassium iodide in about $\frac{1}{5}$ of a test-tubeful of water. Add about 4 cc. of carbon disulfide (or carbon tetrachloride). Shake. Allow to stand. Any color imparted to the carbon disulfide? To this mixture now add a little chlorine water. Shake. Allow to stand. Result?

5. Heat a few small crystals of iodine in a beaker until the iodine vaporizes. Sprinkle some powdered antimony into the beaker. Result?

6. Spread some melted paraffin over both sides of a glass plate. Cut a design through the paraffin on one side with the sharp point of a file or stylus. Place in a lead dish. Cover the glass plate with hydrofluoric acid solution. Allow to stand for about 10 minutes. With a pair of forceps remove the glass plate. Rinse with water and scrape off the paraffin with a knife or melt it off. Result?

Copy the following in your notebook, supplying the word or words needed to make each a complete statement. Underline the words you supply.

Observations and Questions on Experiments

1. The usual delivery tube arrangement for the collection of substances cannot be used with iodine because —?—.

2. In the preparation of iodine, the functions of the reacting substances are: potassium iodide, —?—; sulfuric acid, —?—; manganese dioxide, —?—.

3. Complete the equation: $\text{KI} + \text{H}_2\text{SO}_4 + \text{MnO}_2 \rightarrow$

4. Iodine is evolved in the —?— state. In this state iodine has a —?— color. At room temperature this condenses on the funnel directly to the —?— state.

5. This unusual behavior of iodine is called —?—.

6. An alcohol solution of iodine colors starch —?—.

7. Potassium iodide solution (does, does not) —?— impart a color to the carbon disulfide.

8. After the addition of chlorine to the solution of potassium

iodide, and subsequent shaking the bottom carbon disulfide layer is colored —?— by the free —?— liberated by the chlorine.

9. Complete the equation: $\text{KI} + \text{Cl}_2 \rightarrow$

10. The antimony combines (more, less) —?— vigorously with iodine than with bromine.

11. Complete the equation: $\text{Sb} + \text{I}_2 \rightarrow$

12. Reactions involving the use of hydrofluoric acid are usually carried out in —?— dishes because —?—.

13. The hydrofluoric acid was observed to —?— the glass and —?— the paraffin.

14. Glass is made from sand and is composed of silicates. Complete the following equation: $\text{SiO}_2 + \text{HF} \rightarrow$

Conclusions

1. Iodine is prepared in the laboratory by the reaction of —?—. It is collected by the method of —?—.

2. When heated, iodine is a —?— colored (gas, liquid, solid) —?—. When cooled, this —?— is condensed directly to the —?— state. This process is called —?—.

3. Iodine is (more, less) —?— soluble in water than in solvents such as —?—.

4. Iodine combines (more, less) —?— readily with metals than does bromine.

5. To test for an iodide, add —?— and —?— and shake. A —?— color in the bottom —?— layer indicates that free —?— has been liberated, and that therefore an —?— was present.

Supplementary Exercises

1. Carbon disulfide is not colored by the iodine in —?—, but is colored by —?—.

2. Bromides have a (greater, lesser) —?— heat of formation than iodides and have a —?— heat of formation than chlorides.

3. Bromides are (more, less) —?— stable than iodides and (more, less) —?— stable than chlorides.

4. An important source of iodine is the sodium iodate (NaIO_3) found as an impurity in —?—.

5. Two uses for iodine are: —?—.

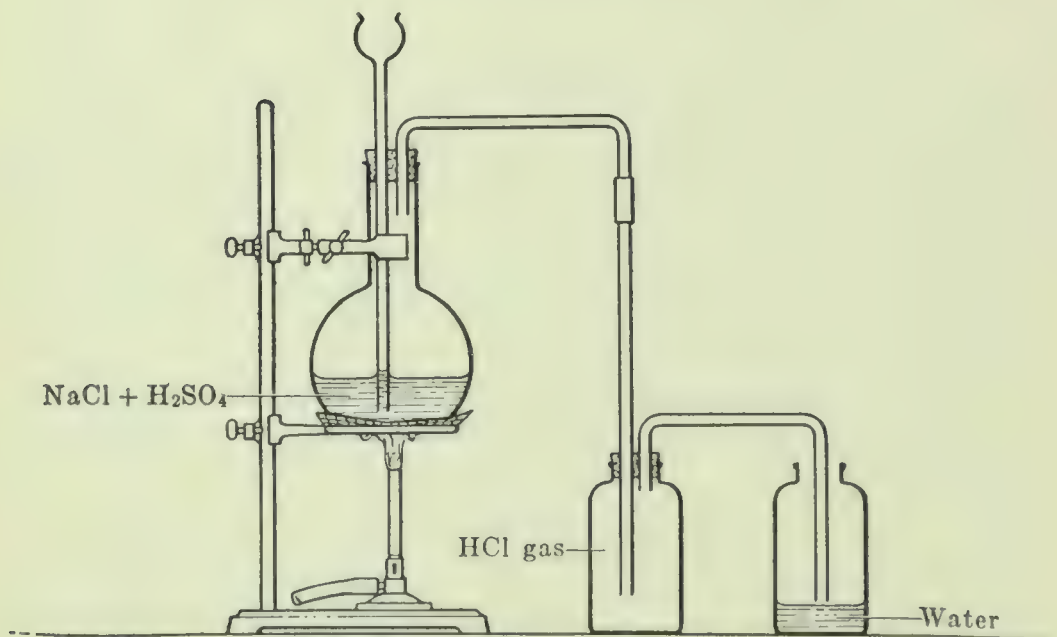
6. Tincture of iodine is a solution of iodine in —?—.
7. Silver iodide is used in —?— because it —?—.
8. The halogen elements arranged in the order of their activity (most active first) are: —?—.
9. Small amounts of iodine compounds seem to be a necessary ingredient of the diet because —?—.
10. Fluorine cannot be prepared by the same laboratory method used for iodine, —?—, and —?— because —?—.
11. In 1886 the French chemist Moissan succeeded in preparing fluorine by electrolysis of a solution of —?— in liquified —?—. The fluorine is collected at the —?— electrode.
12. The most abundant mineral compound of fluorine found in nature is —?—. Its formula is —?—.
13. Write the equation for the preparation of hydrofluoric acid.
14. The reaction in Ex. 13 above is not reversible because —?—.

Optional Questions

1. Compare the physical and chemical properties of chlorine, bromine, and iodine.
2. Which of the halogens will form compounds with the greatest heats of formation? The least? How may we use heats of formation to predict chemical action?
3. Why are the halogen elements grouped together as a chemical family?
4. What special function do sea foods rich in iodine serve in our bodies?
5. What is the true formula of hydrofluoric acid, HF or H_2F_2 ? Explain.
6. What is an insecticide? A germicide? A disinfectant? An antiseptic? Mention a halogen or halogen compound used as each of the foregoing.

UNIT 15. HYDROGEN CHLORIDE AND HYDROCHLORIC ACID

Experiments



Apparatus for Preparing Hydrogen Chloride

1. Arrange apparatus as shown in the figure above. Fill the second collecting bottle to a depth of 1 inch with water and arrange a cardboard cover for it. Be sure to keep the end of the delivery tube just above the surface of the water in this bottle. Put about 2 teaspoonfuls of sodium chloride (17 g.) in the flask and slowly pour about 2 test-tubefuls of sulfuric acid (2:1) through the thistle tube. Heat the mixture gently (and only when necessary) and permit the gas to fill the first bottle and continue into the second bottle containing water. When the first collecting bottle is filled with gas (after a few minutes) remove it; cover with a glass plate; and substitute another collecting bottle. Collect 4 bottles of the gas. The gas should be passed into the water for a total of at least 15 minutes.

2. Collect a test-tubeful of gas. Place your thumb over the mouth and quickly invert, mouth downward, in a dish of water.

Remove your thumb. Allow to stand. Result? Taste a drop of the solution. Result?

3. Blow your breath over the mouth of an open bottle of gas. Result?

4. Bring an open bottle of concentrated ammonium hydroxide (ammonia water) near an open bottle of hydrogen chloride gas. Result?

5. Hold a moistened piece of blue litmus paper over the mouth of an open bottle of gas. Result?

6. Insert a burning wood splint into another bottle of gas. Result?

7. Divide the water solution of this gas in the second collecting bottle into 3 equal portions. Into the first portion drop a strip of magnesium. Into the second portion drop a piece of zinc. Test the gas evolved with burning wood splints. (Hold your thumb over the mouth of the test tube, if necessary, to allow the accumulation of gases.) Results?

8. To the third portion of the water solution add some silver nitrate solution. Result? Expose to light for a few minutes. Result? Test the solubility of one portion of the precipitate with dilute nitric acid. Test the other portion with concentrated ammonium hydroxide. Results? This is a test for a chloride.

Copy the following in your notebook, supplying the word or words needed to make each a complete statement. Underline the words you supply.

Observations and Questions on Experiments

1. The gas prepared and collected in this experiment is called hydrogen chloride. The solution of this gas in water is called hydrochloric acid.

2. This gas is collected by the downward displacement of water method because it is heavier than water.

3. Of the reacting substances in the flask, the purpose of the sodium chloride is to provide chloride ions and the purpose of the sulfuric acid is to provide hydrogen ions.

4. Sodium chloride rather than some other chloride is used because it is soluble in water. Sulfuric acid rather than some other acid is used because it is a strong acid.

5. The bottom of the delivery tube is held just above the surface of the water in the collecting bottle because —?—.

6. Hydrochloric acid is a volatile acid. This means that —?—.

7. When a test tube of hydrogen chloride gas is inverted in a dish of water, the —?—. This shows that hydrogen chloride is —?—.

8. The solution of hydrogen chloride in water has a —?— taste. This is a characteristic of —?—.

9. Blowing your breath over the bottle of hydrogen chloride causes —?—. The —?— of the breath —?— the hydrogen chloride gas to form small liquid particles of —?—.

10. When ammonia gas comes in contact with hydrogen chloride gas, the substance —?— is formed, which appears in the form of —?—. The equation for this reaction is:



11. The moistened blue litmus paper held over the bottle of hydrogen chloride gas turns —?—. This shows the presence of —?—.

12. Hydrogen chloride is (combustible, noncombustible) —?— and (supports, does not support) —?— combustion.

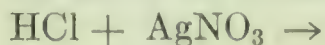
13. Complete the following equations:



14. When hydrochloric acid reacts with a metal, —?— and —?— are formed.

15. A solution of silver nitrate added to hydrochloric acid causes the formation of —?—, which —?— upon exposure to light. This —?— is (soluble, insoluble) —?— in nitric acid and (soluble, insoluble) —?— in ammonium hydroxide.

16. Complete the equations:



Conclusions

1. The equation for the laboratory (and commercial) method of preparing hydrogen chloride is: $\text{NaCl} + \text{H}_2\text{SO}_4 \rightarrow \text{—?—}$, + 2 H₂O

2. This method of preparation for hydrochloric acid is typical

for volatile acids. Concentrated sulfuric acid is used to supply the —?— and because it has a —?— boiling point. The sulfuric acid is made to react with an inexpensive —?— of the acid to be prepared.

3. Three typical properties of acids as illustrated by hydrochloric acid are: Acids have a —?— taste, they change the color of litmus from —?— to —?—, and they react with most metals to form —?— and —?—.

4. To test for the presence of a chloride, add a solution of —?— to the solution of the unknown. A —?— which —?— upon exposure to light, is insoluble in —?— and soluble in —?— indicates the presence of a chloride.

Supplementary Exercises

1. The commercial name of hydrochloric acid is —?—.

2. Two commercial uses for hydrochloric acid are: —?—.

3. In the laboratory preparation of nitric acid (HNO_3), which is also a volatile acid, —?— and —?— are heated together.

4. All acids contain the element —?—. This element is placed first in writing the formula. The rest of the acid is called the —?— radical.

5. Vinegar contains —?— acid; oranges, lemons, and cranberries contain —?— acid; sour milk contains —?— acid; rancid butter contains —?— acid.

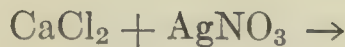
6. It can be shown that a substance contains an acid by testing with —?—.

7. Three metals that will not liberate hydrogen from acids are: —?—. These metals are (more, less) —?— active than hydrogen.

8. The gastric juice in the stomach contains a little hydrochloric acid. Its purpose is —?—.

9. "Aqua regia" is a mixture of —?— and —?— and is used to —?—.

10. Complete the equations:



Optional Questions

1. An unlabeled bottle was suspected to contain a solution of hydrochloric acid. How could you prove this chemically?
2. State the names and formulas of the chlorine series of acids. What are the names and formulas of the sodium salts of these acids?
3. State the names and formulas of the halogen acids similar to hydrochloric acid.
4. Write balanced equations for the action of hydrochloric acid with zinc oxide, ferric oxide, sodium hydroxide, and calcium hydroxide.
5. Why cannot hydriodic acid be made by the same method as hydrochloric acid? How is hydriodic acid prepared?
6. Compare the stability of the four halogen acids. Compare their relative reducing power.
7. Describe an experiment which demonstrates the great solubility of hydrogen chloride in water.
8. What is a salt? Give the names and formulas of 4 salts of hydrochloric acid; of sulfuric acid; of nitric acid.

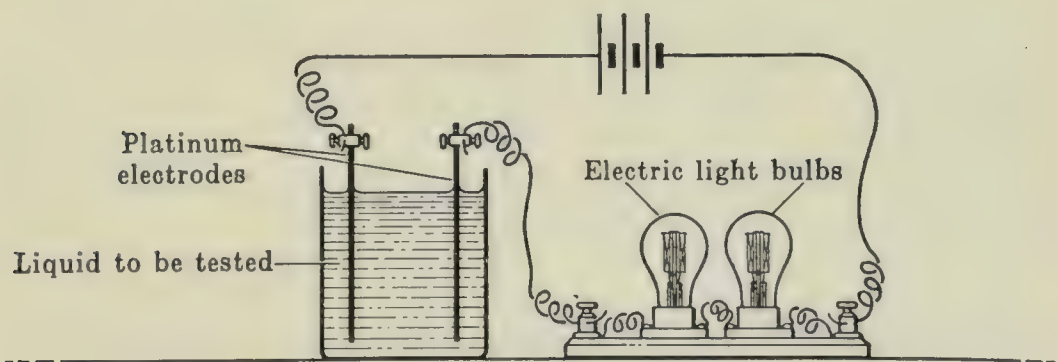
UNIT 16. IONIZATION AND THE ELECTRON THEORY

Experiments

NOTE: *It is suggested that the instructor demonstrate at least Exp. 3.*

1. Place two clean, *dry* glass plates on the table. On the first place a small quantity of tartaric acid, on the second a small quantity of slaked lime (calcium hydroxide). Place a piece of blue litmus paper on the tartaric acid and a piece of red litmus paper on the slaked lime. Any result? Now carefully place one drop of water on each substance. Do the moistened substances affect the litmus?

2. Pulverize a few crystals of copper bromide on a piece of white paper. What is the resulting color? Dissolve a little of this copper bromide in methyl alcohol (or ethyl alcohol). What is the resulting color? Then dissolve a little in a few drops of water. Color? Add more water. Any change in color? Dissolve some anhydrous white copper sulfate in water. Any color change?



Apparatus for Determining Conductivity

3. (a) Arrange apparatus as shown in the figure above. (Use two 40-watt lamps, if the house current is used. If a storage battery or 4 dry cells in series are employed as the source of current, use a 6-volt lamp.) The electric light bulbs in series with the electrodes will indicate whether or not electricity is passing through the various substances to be tested. The brightness of the lamp will indicate the degree of conductivity. Use either

platinum or carbon electrodes. For the containers, use either beakers or glass tumblers.

(b) Try to pass electricity through distilled water. Result? Then clean and wipe dry the electrodes and try to pass electricity through dry sodium chloride. Result? While the electrodes are immersed in the sodium chloride, pour a little distilled water into the container. Result? Pour more water into the container. Result?

(c) Try to pass the electric current through 10% solutions of all the substances listed in the table below. Wash the electrodes thoroughly with distilled water after each test. Make a table like the one below in your notebook and record your results in it.

TABLE OF CONDUCTIVITY

SUBSTANCE USED	BRIGHTNESS OF LAMP (DIM, BRIGHT, NO LIGHT)	GOOD ELECTROLYTE, POOR ELECTROLYTE, NON-ELECTROLYTE	SUBSTANCE USED	BRIGHTNESS OF LAMP (DIM, BRIGHT, NO LIGHT)	GOOD ELECTROLYTE, POOR ELECTROLYTE, NON-ELECTROLYTE
Distilled water		Poor	NaOH solution		Good
Dry NaCl (solid)		Poor	NH ₄ OH solution		Good
NaCl solution		Good	CuSO ₄ solution		
HCl solution		Good	NaC ₂ H ₃ O ₂ solution		Poor
HNO ₃ solution		Good	KNO ₃ solution		Good
H ₂ SO ₄ solution			Glycerine solution		Poor
HC ₂ H ₃ O ₂ solution		Poor	Sugar solution		Poor
KOH solution		Good	Alcohol solution		Poor

Copy the following in your notebook, supplying the word or words needed to make each a complete statement. Underline the words you supply.

Observations and Questions on Experiments

1. Litmus is changed from —? —to —?— by (wet, dry, either wet or dry) —?— tartaric acid.

2. Litmus is changed from —?— to —?— by (wet, dry, either wet or dry) —?— calcium hydroxide.

3. Copper bromide is colored —?— in the dry state. It is colored —?— in methyl alcohol solution, —?— in concentrated water solution, and —? —in dilute water solution.

4. Copper sulfate is colored —?— in the dry (anhydrous) state. In water solution it is colored —?—. This is used as a simple test for water.

5. Platinum or carbon electrodes are used because —?—.

6. Distilled water (did, did not) —?— permit the passage of electricity.

7. Dry sodium chloride (did, did not) —?— permit the passage of electricity.

8. By adding water to the dry sodium chloride the effect on the lamp was —?—, indicating that —?— was passing through the entire circuit.

9. Those substances whose water solutions conduct electricity are called —?—. They conduct electricity because they form —?— when dissolved in water.

10. Substances which conduct electricity (are alike, vary) —?— in their conducting power because —?—.

11. The substances which do not conduct the electric current are called —?—.

Conclusions

1. Three classes of compounds that are electrolytes are: —?—.

2. Electrolytes in water solution conduct electricity because —?—.

3. Acids and bases do not exhibit their typical properties

unless dissolved in —?—. The typical properties of acids are attributed to the —?— ions and of bases to the —?— ions that they form in solution.

4. The —?— color of copper compounds in water solution is due to the —?— ions.

5. The properties of ions (are the same as, differ from) —?— the properties of the corresponding atoms.

Supplementary Exercises

1. Electrolytes are (polar, nonpolar) —?— compounds and are formed by the borrowing and lending of —?—.

2. Nonelectrolytes are (polar, nonpolar) —?— compounds and are formed by the sharing of —?—.

3. A metallic atom is changed to a (+, —) —?— electrically charged ion by the (gain, loss) —?— of electrons.

4. A nonmetallic atom is changed to a (+, —) —?— electrically charged ion by the (gain, loss) —?— of electrons.

5. Many electrolytes will ionize in the absence of water or other solvents provided that they are in —?— state.

6. The number of + or — electrical charges on any ion is dependent upon —?—.

7. The active acids and bases and most salts are known as strong electrolytes because they ionize —?— in dilute solutions. The slightly active acids and bases are known as weak electrolytes because they ionize —?— even in dilute solutions.

8. Copy and complete the ionic equations for the dissociation in water of sodium chloride, sulfuric acid, potassium hydroxide, and ferric chloride.



9. While electricity is passing through the solution of an electrolyte, the metallic or + ions are attracted to the —?— and the nonmetallic or — ions are attracted to the —?— thus causing the —?— of the electrolyte.

10. The degree of dissociation (or amount of ionization) of an electrolyte depends upon: —?—.

11. Substances in water solution usually react more readily than in the solid state because —?—.

12. Dissolved polar compounds lower the freezing point and raise the boiling point of water (less, more) —?— than dissolved nonpolar compounds.

13. The electrolytic dissociation theory, otherwise known as the theory of ionization, was first advanced in 1887 by the Swedish chemist —?—.

14. According to the electron theory all atoms are composed of —?—. Elements differ from each other because of different arrangements and numbers of —?— within their atoms.

15. The nucleus of an atom of any element contains —?— and —?—.

16. The law of atomic numbers states that the 92 elements can be arranged in a table starting with —?— having an atomic number of —?— and ending with the heaviest atom, that of —?—, having an atomic number of —?—.

17. The atomic number of an element is numerically equal to the number of free —?— in the nucleus.

18. The atomic number of an element is also equal to the number of —?— electrons.

19. The mass (weight) of an atom is concentrated in its —?—.

20. A hydrogen atom weighs approximately —?— times as much as an electron.

21. An atom may be represented graphically by arranging the —?— in successive rings around the nucleus. The first complete ring contains —?—; the second complete ring contains —?—; the third complete ring contains —?—; and the fourth complete ring contains —?—.

22. The valence of an element is the number of —?— which its atom must lend or borrow to complete its outermost ring.

23. Metallic elements contain —?— the number of electrons necessary to complete the outermost ring. Metals are (borrowers, lenders) —?— of electrons.

24. Nonmetallic elements contain —?— the number of elec-

trons necessary to complete the outermost ring. Nonmetals are (borrowers, lenders) —?— of electrons.

25. An element is amphoteric if the outermost ring contains —?— the number of electrons necessary to complete the ring.

26. The chemical activity of an atom depends upon the —?— in its outermost ring. When the outermost ring is complete the element is —?— and has a valence of —?—.

27. Chemical union between atoms is the —?— or —?— of the electrons in the outermost rings until a stable condition is reached.

28. Isotopes are elements whose atoms have the same —?— but differ from each other in —?—.

29. Draw labelled diagrams to show the probable structure of the following atoms: hydrogen, oxygen, chlorine, aluminum, sodium, magnesium, sulfur, argon, nitrogen.

Optional Questions

1. Polar compounds or electrolytes lower the freezing point of water much more than nonpolar compounds. Yet alcohol and glycerine are preferred in automobile radiators. Why?

2. Explain the ionization of sodium chloride in water in terms of the electron theory.

3. Explain what happens at the electrodes when an electric current is sent through a solution of sodium chloride.

4. Explain with the aid of a diagram how a spoon can be electroplated with silver.

5. Explain how the electron and ionization theories support each other.

6. Explain, from the electron theory point of view, what happens when an electric current is sent through a solution of copper sulfate.

7. Explain the recent modifications of the ionization theory with respect to dissociation of strong electrolytes and the hydration of ions.

8. Who discovered the electron? Describe the experiments by which he determined the nature of the electron.

9. By means of diagrams show what happens when atoms of magnesium and chlorine unite. When sodium and oxygen unite. When calcium and sulfur unite.

10. An atom has a nucleus of 20 protons and 20 neutrons. Make a diagram of the atom and describe some of its properties.

11. Who discovered X-rays? How are they produced? What are they capable of doing?

12. Who discovered radioactivity? How did this discovery lead to the discovery of radium? Give a brief account of the work leading up to the isolation of radium.

13. Discuss the remarkable properties of radium. What three types of rays does it give off?

14. "The conception of the structure of the atom makes it possible for the present-day scientist to explain the riddle of transmutation." Explain this statement.

UNIT 17. OXIDATION AND REDUCTION

Experiments

1. Hold a small piece of magnesium in a pair of forceps and ignite it. Result?

2. Carve out a small cavity in a block of solid carbon dioxide (dry ice). Place in this cavity some powdered magnesium and mix it with some powdered solid carbon dioxide scraped from the block. Stick a two inch strip of magnesium ribbon into the mixture. Ignite. **Stand a few feet away.** Result?

3. Strongly heat a mixture of a little powdered copper oxide with about an equal volume of powdered charcoal or coke in a hard glass test tube for about 10 minutes. Pass the gas evolved into a test tube half filled with limewater. Result? What gas is evolved? Stop heating. Examine the contents of the test tube on a sheet of white paper. What do you observe? The copper oxide has undergone the process of reduction.

4. To a test tube $\frac{1}{3}$ full of freshly prepared ferrous chloride solution add a few drops of concentrated hydrochloric acid. Then add 5 cc. of hydrogen peroxide. Boil for 2 or 3 minutes. Test some of the resulting mixture for the presence of a ferric salt (ferric chloride). Result?

NOTE: *To test for a ferric salt add a few drops of a solution of potassium ferrocyanide. A dark blue precipitate called Prussian blue will indicate the presence of a ferric salt. Another more sensitive test is to add a few drops of potassium sulfocyanate. A dark red coloration will indicate the presence of a ferric salt.*

5. To a test tube $\frac{1}{3}$ full of ferric chloride solution add a few drops of concentrated hydrochloric acid and a small quantity of iron filings. Heat gently. There should be a vigorous evolution of gas. What gas? After a few minutes, test a small portion of the solution for the presence of a ferrous salt (ferrous chloride). Result?

NOTE: *To test for a ferrous salt add a few drops of a solution of*

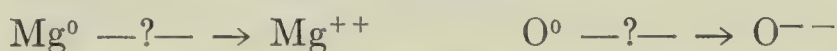
potassium ferricyanide. A dark blue precipitate called Turnbull's blue will indicate the presence of a ferrous salt.

Copy the following in your notebook, supplying the word or words needed to make each a complete statement. Underline the words you supply.

Observations and Questions on Experiments

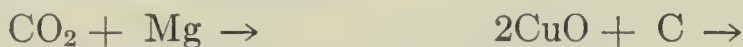
1. When magnesium combines with oxygen, the process is called —?—. The equation is: —?—.

2. Complete the electronic equations:



3. In the reaction between magnesium and oxygen, the magnesium (loses, gains) —?— electrons and the oxygen (loses, gains) —?— electrons. The magnesium is said to undergo the process of (oxidation, reduction) —?— and the oxygen is said to undergo the process of (oxidation, reduction) —?—.

4. Complete the equations:



In these reactions the carbon dioxide and the copper oxide are said to be (oxidized, reduced) —?—; the magnesium and carbon are the —?— agents.

5. Solid carbon dioxide is at a temperature of approximately —?— degrees C. or —?— degrees F. When magnesium reacts with solid carbon dioxide, the temperature changes are —?—.

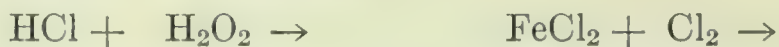
6. Carbon is (more, less) —?— active than magnesium and (more, less) —?— active than copper.

7. Complete the electronic equations:



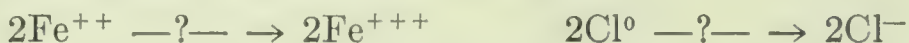
8. In the reaction between copper oxide and carbon, the copper ions (lose, gain) —?— electrons and become copper —?—. The copper is undergoing the process of (oxidation, reduction) —?—. The carbon atoms (lose, gain) —?— electrons and become carbon —?—. The carbon atoms are undergoing the process of (oxidation, reduction) —?—.

9. Complete the equations:



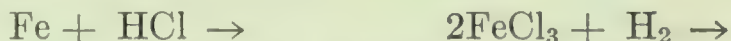
10. When ferrous chloride is converted into ferric chloride, the valence of iron is (lowered, raised) —?— from —?— to —?—.

11. Complete the electronic equations:



12. In converting a ferrous salt to a ferric salt, the ferrous ions (lose, gain) —?— electrons to form ferric ions. In this reaction the (lowering, raising) —?— of the valence and the (loss, gain) —?— of electrons by the iron is called (oxidation, reduction) —?—. The (loss, gain) —?— of electrons by the chlorine atoms is called (oxidation, reduction) —?—.

13. Complete the equations:



14. When ferric chloride is converted to ferrous chloride, the valence of iron is (lowered, raised) —?— from —?— to —?—.

15. Complete the electronic equations:



16. In converting a ferric salt to a ferrous salt, the ferric ions (lose, gain) —?— electrons to form ferrous ions. In the reaction the (lowering, raising) —?— of the valence and the (loss, gain) —?— of electrons by the iron is called (oxidation, reduction) —?—. The (loss, gain) —?— of electrons by the hydrogen atoms is called (oxidation, reduction) —?—.

Conclusions

1. Oxidation may be considered as being:

- (a) The combination of a substance with —?—.
- (b) The —?— of valence.
- (c) The (loss, gain) —?— of electrons.

2. Reduction may be considered as being:

- (a) The —?— of oxygen from a substance.
- (b) The —?— of valence.
- (c) The (loss, gain) —?— of electrons.

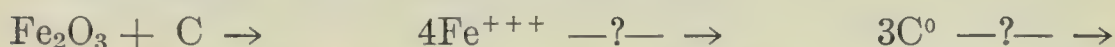
3. Oxidation and reduction reactions (may occur independently of each other, always occur together) —?—.

Supplementary Exercises

1. Four oxidizing agents are: —?—.

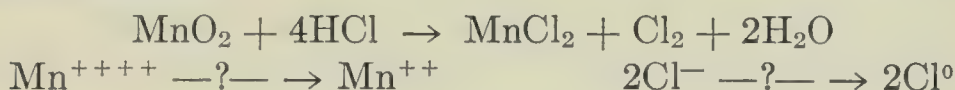
2. Four reducing agents are: —?—.

3. Complete the following formula and electronic equations for the action of ferric oxide with carbon.



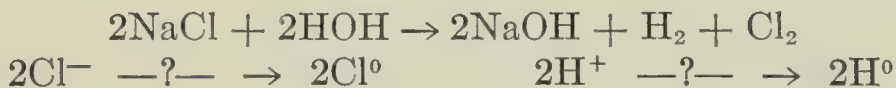
4. In the reaction between ferric oxide and carbon, the —?— is the oxidizing agent, the —?— is the reducing agent, the —?— is undergoing the process of oxidation, and the —?— is undergoing the process of reduction.

5. Complete the electronic equations for the oxidation-reduction reaction:



6. In the reaction between manganese dioxide and hydrochloric acid, the —?— is the oxidizing agent, the —?— is the reducing agent, the —?— is undergoing the process of oxidation, and the —?— is undergoing the process of reduction.

7. Complete the electronic equations for the electrolysis of brine:



8. In the electrolysis of brine reaction, the —?— ions are attracted to the anode. At the anode they (lose, gain) —?— electrons and are therefore undergoing the process of (oxidation, reduction) —?—. The —?— ions are attracted to the cathode. At the cathode they (lose, gain) —?— electrons and are therefore undergoing the process of (oxidation, reduction) —?—.

9. In all electrolysis reactions, the cathode always (gives, removes) —?— electrons and therefore causes (oxidation, reduc-

UNIT 18. ACIDS, BASES, AND SALTS

Experiments

1. Acids. To a test tube $\frac{1}{2}$ full of water add just a few drops of hydrochloric acid. Stir with a clean stirring rod. Touch the rod to pieces of red and blue litmus paper, placed on a clean glass plate. Result? Taste by touching the rod to your tongue. Result? Repeat using sulfuric, nitric, and acetic acids.

Copy and complete the following table.

ACID	FORMULA	TASTE	COLOR CHANGES WITH LITMUS	IONS IN SOLUTION
Hydrochloric acid				H ⁺ and Cl ⁻
Sulfuric acid				
Nitric acid				
Acetic acid				

2. Bases. Repeat Exp. 1 above using solutions of the following bases: sodium hydroxide, potassium hydroxide, calcium hydroxide (use full strength), and ammonium hydroxide. Results?

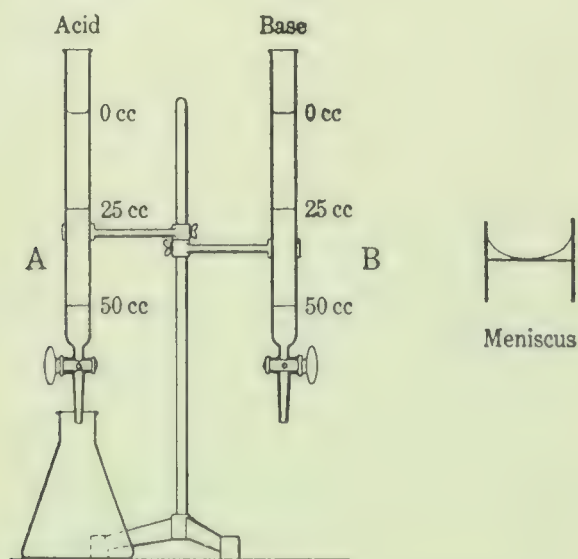
Copy and complete the following table.

BASE	FORMULA	TASTE	COLOR CHANGE WITH LITMUS	IONS IN SOLUTION
Sodium hydroxide				Na ⁺ and OH ⁻
Potassium hydroxide				
Calcium hydroxide				
Ammonium hydroxide				

3. Neutralization. Have a test tube $\frac{2}{3}$ full of dilute hydrochloric acid (1:4) and another test tube $\frac{2}{3}$ full of dilute sodium hydroxide solution (10%) ready in a test tube rack. Put strips of blue and red litmus paper on a clean glass plate. Pour $\frac{1}{2}$ of a test-tubeful of the sodium hydroxide solution into a clean evapo-

rating dish. Slowly add $\frac{1}{2}$ of a test-tubeful of the hydrochloric acid solution to the sodium hydroxide in the dish. Stir with a clean stirring rod. Bring a drop of the mixture at the end of the stirring rod to each kind of litmus paper on the glass plate. If the effect is acid, slowly add more base (if the effect is basic slowly add more acid) a drop at a time and keep stirring and testing as before with litmus paper until the mixture is neutral (*i. e.*, has no effect on litmus paper). Taste *a drop* of the solution. Result?

Place the evaporating dish on a wire gauze supported on a ring stand and evaporate the solution to dryness. Examine and taste the residue.



Apparatus Used in Titration

4. Neutralization.

(Optional) Titration.

Set up two clean burettes as shown in the diagram. Fill burette A with a dilute solution of hydrochloric acid (one part concentrated hydrochloric acid to ten parts of water). Fill burette B with a molar solution of sodium hydroxide. This molar solution contains *exactly* one molecular

weight in grams (40 grams) of sodium hydroxide in one liter of solution. If the sodium hydroxide is not 100% pure, allowance for its purity should be made in making up the molar solution; e.g., if 95% pure use 42 grams in making one liter of solution.

The burettes are filled to the "0" (zero) point exactly when the bottom of the meniscus coincides with the "0" line. See the diagram above. This is done by filling the burettes somewhat above the "0" mark and letting the liquid run out slowly until the "0" mark is reached.

Allow about 10 cc. of the acid solution to drop into a small, clean Erlenmeyer flask (or a beaker). The exact volume can be determined by taking the two readings indicated in the Table. Place a sheet of white paper under the flask containing the acid solution. Add two drops of phenolphthalein solution. (See note below.) Slowly drop sodium hydroxide solution from the other burette into this flask, rotating or stirring the flask constantly. When a pink color tends to persist as the base comes in contact with the indicator, add the basic solution *slowly*, drop by drop. When a single drop of the basic solution causes the entire liquid in the flask to change *and remain* a faint pink, the neutralization is complete. This condition is known as the *end point* of the titration.

Record the volumes of solution used in the Table on page 86. Repeat the experiment to get a second set of readings. It is not necessary to start at the "0" point. The volume of either the acid solution or the basic solution can be obtained by taking two readings (one *before* and one *after*) and subtracting.

Calculate the weight of hydrogen chloride in 1 liter of the solution of hydrochloric acid used.

NOTE: *Phenolphthalein is a white, crystalline, weak organic acid which is colorless in an acid or a neutral solution. The addition of a base, or alkali, in slight excess produces a new substance. This new substance is pink in color. Such substances as litmus and phenolphthalein are called indicators. The indicator solution used in this experiment is made by dissolving 1 gram of phenolphthalein in 100 cc. of alcohol.*

	FIRST TITRATION	SECOND TITRATION
Final reading of acid	cc.	cc.
Initial reading of acid	cc.	cc.
Volume of acid used	cc.	cc.
Final reading of base	cc.	cc.
Initial reading of base	cc.	cc.
Volume of base used	cc.	cc.
Weight of base in 1cc. of solution	0.04 gm.	0.04 gm.
Weight of base used	gm.	gm.
Calculated weight of acid to be used	gm.	gm.
Calculated weight of acid in 1cc. of solution	gm.	gm.
Calculated weight of acid in 1 liter of solution	gm.	gm.

1 liter (1000 cc.) of a molar solution of sodium hydroxide contains 40 grams of sodium hydroxide.

1 cc. of a molar solution of sodium hydroxide contains 0.04 gram of sodium hydroxide.

Weight of
NaOH used

Weight of
HCl used



23

1

16

35.5

1

36.5

40

$$\frac{\text{Weight of NaOH}}{40} = \frac{\text{Weight of HCl}}{36.5}$$

Calculated weight of acid in 1 cc. of hydrochloric acid solution = $\frac{\text{Weight of HCl}}{\text{No. of cc. of HCl used.}}$

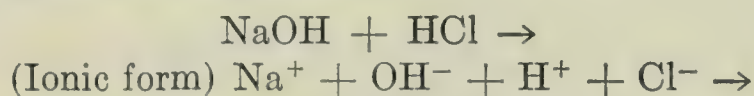
Copy each of the following in your notebook, supplying the word or words needed to make each a complete statement. Underline the words you supply.

Observations and Questions on Experiments

1. Sodium hydroxide solution has a —?— taste, and turns litmus from —?— to —?—. Hydrochloric acid solution has a —?— taste, and turns litmus from —?— to —?—. When these two are mixed in equivalent proportions, the mixture has a —?— taste, and the effect on the litmus is —?—.

2. When an acid and a base react, the reaction is called a —?— reaction.

3. Complete the equations:



4. When a neutral mixture of NaOH and HCl is evaporated to dryness, the —?— is driven off and the residue tastes —?—. The residue is the substance —?—.

Conclusions

1. Water solutions of acids contain —?— ions, change litmus from —?— to —?—, and have a —?— taste. Similarities in the behavior of acids are attributable to the —?— ions which they possess in common.

2. Water solutions of bases contain —?— ions, change litmus from —?— to —?—, and have a —?— taste. Similarities in the behavior of bases are attributable to the —?— ions which they possess in common.

3. When an acid reacts with a base (—?— reaction), the characteristic —?— ions of the acid combine with the characteristic —?— ions of the base to form neutral —?—. The —?— of the base combines with the —?— of the acid to form a salt.

Supplementary Exercises

1. Acids usually react with metals to form —?— and —?—.
2. Water solutions of bases usually have a —?— feel.

3. Not every substance that contains hydrogen is an acid. The hydrogen of the substance must form —?— when in water solution to be acidic hydrogen.

4. Some acids found in the home are: —?— acid in vinegar; —?— acid in grapes; —?— acid in lemons, oranges and cranberries; —?— acid in rancid butter; oxalic acid used for —?—; and boric acid used for —?—.

5. Housekeepers store pickles in glass or earthenware containers rather than in metal ones because —?—.

6. Some acids do not change blue litmus to red because —?—.

7. Caustic soda is the common name for —?— and is so called because —?—. It is also called "lye" which is the abbreviation of the word —?—.

8. Two ways of preparing bases are illustrated by the following equations. Copy and complete:



9. An alkali is a substance whose water solution changes litmus from —?— to —?—.

10. Alkalies dissolve —?— and are therefore useful in cleaning.

11. Some metallic hydroxides, such as ferric hydroxide, do not exhibit the typical basic properties because —?—.

12. Neutralization reactions are (endothermic, exothermic) —?—.

13. "Acid mouth" is neutralized by dentifrices containing —?— substances such as —?— and —?—.

14. If you spill some acid on your clothing, you should neutralize it with a —?— such as —?— which does not —?—.

15. Indigestion is usually associated with (an acid, an alkaline) —?— condition of the stomach. To relieve this condition, mild (acids, alkalies) —?— such as —?— and —?— are taken.

16. Complete the following equations and then write them in ionic form:



Optional Questions

1. Starch and sugar contain hydrogen. Are they acids? Explain.
2. What is an acid radical? Explain, giving illustrations.
3. What is a salt? Describe 4 different ways of preparing a salt, using sodium chloride as an example.
4. What is meant by the heat of neutralization? This heat is expressed in calories. What is a calorie?
5. Explain why bottles of sodium hydroxide solution are equipped with rubber stoppers.
6. How does a farmer test his soil for excess acidity or alkalinity? What are his remedies for these conditions?
7. What is a monobasic acid? A dibasic acid? Illustrate.
8. From the point of view of the ionization theory, explain what happens when:
 - (a) Sodium hydroxide is dissolved in water.
 - (b) Sulfuric acid is dissolved in water.
 - (c) Sodium hydroxide and sulfuric acid are mixed in equivalent proportions.
 - (d) The solution resulting from (c) is evaporated to dryness.
9. Why are sodium hydroxide and potassium hydroxide solutions used in testing textile fabrics?
10. What is meant by the pH value of a solution? What is the pH value of pure water? What pH values indicate acidity? alkalinity?
11. Describe the commercial process for making sodium hydroxide. State three uses for sodium hydroxide and for each of the byproducts in this process.
12. In the form of a table indicate the chemical name, common name (if any), formula, and two uses for each of five different sodium salts.
13. Discuss recent changes in the interpretation of the theory of ionization with respect to acids and bases and involving these terms: oxonium ion, proton donor, proton acceptor.

UNIT 19. SOLUTIONS, SUSPENSIONS, AND COLLOIDAL SUSPENSIONS

Experiments

1. Place a little salt, sand, starch, and one or two small crystals of potassium permanganate, respectively, in 4 wide mouth bottles. Fill the bottles with water. Place your hand over the mouth of each bottle and shake for a minute or two. Set the bottles aside for about 15 minutes. (In the meantime do Exp. 2.) Examine each bottle. Are the particles visible? Have they settled upon standing? Is the color of the potassium permanganate mixture uniform throughout? Taste the top layer of the salt mixture. Then pour out some of the contents and taste the remaining part. (Leave some for Exp. 3.) Is there any difference in taste?

2. Place a **very small** quantity of starch in a test tube. Fill the test tube with water and shake thoroughly. Heat the mixture gently (do not boil). Compare this mixture of starch and hot water (colloidal suspension) with the mixture of starch and cold water in Exp. 1 with respect to size, visibility, and settling of particles.

3. Filter a little of each mixture of Exps. 1 and 2 into test tubes. (This work should be distributed throughout the class or demonstrated by the instructor and the result observed by everybody.) Which ones pass through the filter paper? How do you know? Test both starch filtrates with tincture of iodine. Results?

Copy the following in your notebook, supplying the word or words needed to make each a complete statement. Underline the words you supply.

Observations and Questions on Experiments

1. After mixing thoroughly with water, the particles of —?— and —?— are visible and the particles of —?— and —?— are not visible.

2. After mixing thoroughly with water, the particles of —?— and —?— do not settle upon standing and the particles of —?— and —?— settle to the bottom.

3. After mixing thoroughly with water, the particles of —?— and —?— are distributed equally or uniformly throughout the water. In the case of —?— this was shown by —?— and in the case of —?— this was shown by —?—.

4. —?— and —?— are said to be in solution in water. They are called solutes and the water is called the —?—.

5. —?— and —?— are said to be in suspension in water.

6. The starch particles in hot water differ from the starch particles in cold water in these three particulars: —?—.

7. The starch in —?— water is in suspension.

8. The starch in —?— water is in colloidal suspension.

9. The particles of the following three substances pass through the filter paper and are found in the filtrate: —?—.

10. The particles of the following substances do not pass through the filter paper and are found on the filter paper: —?—.

11. Starch is colored —?— by iodine. This is used as a test either for starch or for iodine.

Conclusions

1. A substance is said to be in solution when its particles are: —?—.

2. The substance that goes into solution is called the —?—.

3. The substance that does the dissolving is called the —?—.

4. A substance is said to be in suspension when its particles are: ~~evenly distributed~~ *evenly distributed*.

5. Colloidal particles are (smaller, larger) —?— than particles in solution and are (smaller, larger) —?— than particles in suspension.

6. Particles in —?— and —?— pass through ordinary filter paper. Particles in —?— do not.

Supplementary Exercises

1. Colloidal particles can be filtered out through —?—.

2. Ordinary filtration will rid water of —?— impurities.

3. A solution is a (mixture, compound) —?—; it (conforms, does not conform) —?— to the law of definite proportions.

4. A dilute solution contains —?— amount of solute.

5. A concentrated solution contains —?— amount of solute.

6. A saturated solution contains —?—.
7. A supersaturated solution contains —?—.
8. Two liquids which dissolve in each other in all proportions are —?—. These liquids are said to be —?—.
9. Two nonmiscible liquids when shaken together vigorously form —?—.
10. It is impossible to make a concentrated solution of lime-water because —?—.
11. Three well known solvents other than water are: —?—.
12. —?— is a noninflammable organic solvent very useful in cleaning because it easily dissolves —?—.
13. Pure water will boil at (a lower, a higher, the same) —?— temperature and will freeze at (a lower, a higher, the same) —?— temperature than water containing a dissolved solute such as salt.

Optional Questions

1. Name some mineral substances usually found dissolved in water. Name some mineral substances usually found suspended in water.
2. Why is it impossible to find chemically pure water in nature?
3. What is tincture of iodine? Why is not water used as the solvent?
4. How would you determine experimentally whether a solution was dilute? concentrated? saturated? unsaturated? supersaturated?
5. Explain how the solution of a given substance in water may be hastened.
6. What is the difference between a crystalloid and a colloid?
7. What is meant by dialysis?
8. Explain the meaning of "gelation" and "peptization" as applied to colloids.
9. Cite examples of the following solutions: gas in a liquid, liquid in a liquid, gas in a solid, liquid in a solid, solid in a gas, solid in a solid.
10. What is meant by the statement that the air in a room has a relative humidity of 40%?
11. Why is gasoline used for cleaning purposes?
12. Compare the approximate sizes in millimeters (mm.) of particles in suspension, in colloidal suspension, and in true solution.
13. What is a molar solution? A normal solution?

UNIT 20. AIDS TO SOLUTION

Experiments

NOTE: *Do not throw away the solutions of copper sulfate and potassium nitrate that you make in this experiment. Place them in receptacles provided by the instructor.*

1. Take 2 small crystals of copper sulfate (blue vitriol) of about the same size. Pulverize one of the crystals, using a mortar and pestle. Put the powdered copper sulfate in one test tube and the crystal in another test tube. Fill each test tube $\frac{2}{3}$ full of water. Place a thumb over the mouth of each test tube and shake. Which dissolves first, the crystal or the powder?

2. Pulverize another crystal of copper sulfate of about the same size. Place the powder in a test tube and fill $\frac{2}{3}$ full of water. This time allow the test tube to stand without shaking. Does shaking aid solution?

3. To a test tube filled to a depth of one inch with cold water add sufficient potassium nitrate (saltpeter) to form a saturated solution. Shake or use a stirring rod to aid solution throughout this experiment. Leave a few crystals of potassium nitrate on the bottom of the test tube. Heat the solution gently for a minute. Any result? Add more potassium nitrate until the slightly warmed water is again saturated, again leaving a few crystals on the bottom of the test tube. Heat again. Any result?

4. Fill a test tube to a depth of 1 inch with very dilute ammonium hydroxide (solution of ammonia gas in water). Heat gently. Smell the gas evolved. Hold a piece of moist red litmus paper over the mouth of the test tube. Results? How does heating affect the solubility of a gas?

Observations and Questions on Experiments

1. The powdered copper sulfate dissolves (as fast as, faster than, slower than) —?— the crystal of copper sulfate.

2. The powdered copper sulfate has (more surface, less surface, as much surface) —?— exposed to the solvent action of water than the crystal of copper sulfate.

3. Shaking or stirring the copper sulfate in water will (hasten, slow up, have no effect on) —?— the formation of the solution.

4. Potassium nitrate is (more soluble, less soluble, equally soluble) —?— in hot water than in cold water.

5. A saturated solution of potassium nitrate at 60° C. would contain (more, less, the same amount) —?— of potassium nitrate than a saturated solution at 20° C.

6. Heating the solution of ammonia gas in water has the effect of making the gas —?—.

7. In order to aid the solution of ammonia gas in water we should use (hot water, cold water, water at room temperature) —?—.

Conclusions

1. Three ways of aiding (increasing the speed of) the solution of solids in liquid solvents are: —?—.

2. With a rise in temperature, a liquid solvent will usually dissolve (more, less, the same amount) —?— of solid solutes and —?— of gaseous solutes.

Supplementary Exercises

1. Increased pressure will aid the solubility of —?— and will have no effect on the solubility of —?—.

2. A commonly used household substance that is only slightly more soluble with a rise in the temperature of the solvent is —?—.

3. When a hot saturated solution of a solid solute is allowed to cool, —?— usually form.

4. A glass of water taken from the faucet will, upon standing for some time, usually evolve bubbles of —?— because —?—.

5. In making a boric acid solution for use as an antiseptic, the boric acid is always dissolved in —?— water because —?—.

6. In making iced tea, the tea is first placed in (hot, cold) —?— water because —?—.

7. A bottle of soda pop foams when it is first opened because —?—.

8. Water that has first been heated produces clear ice when frozen because —?—.

9. If some copper sulfate were suspended in a piece of cheese cloth and hung near the top surface of some water, the force of —?— would aid solution by —?—.

Optional Questions

1. When sea water is evaporated, why does the sodium chloride separate out before the other minerals?

2. How would you prepare a supersaturated solution of "hypo"?

3. What gas is dissolved in water to make vichy or soda water? In making vichy or soda water, what conditions of temperature and pressure are necessary?

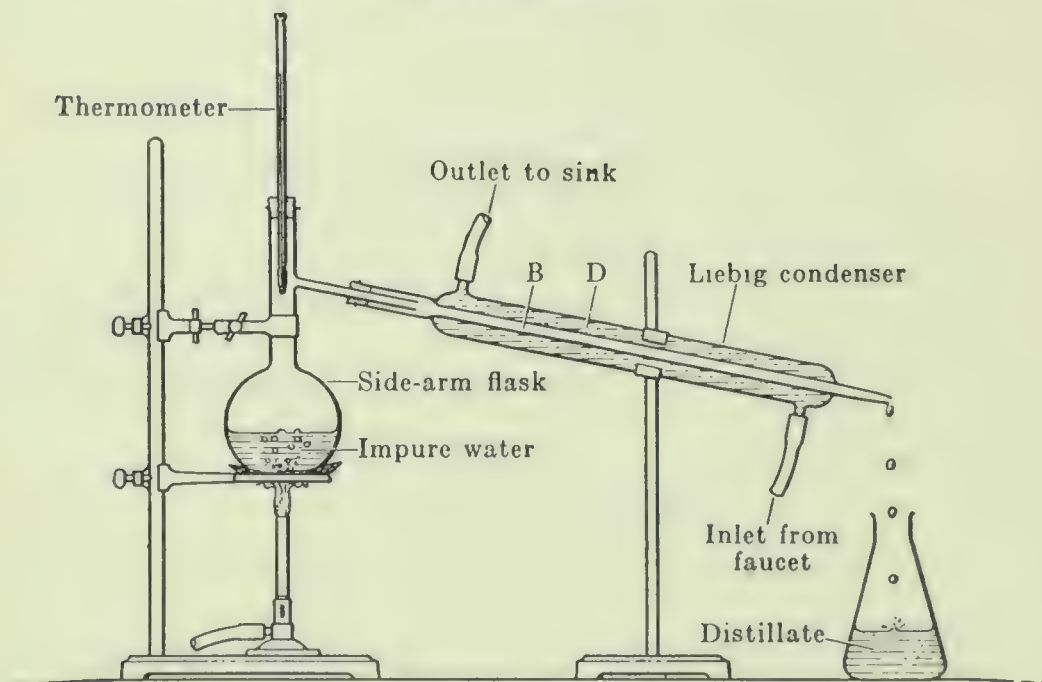
4. From the point of view of the kinetic theory of matter explain the various methods of speeding up solution.

5. What are protective colloids? Describe the action of a protective colloid.

6. What would you add to a mixture of turpentine and alcohol, which are normally nonmiscible, to help make the mixture permanent?

UNIT 21. DISTILLATION OF WATER

Experiments



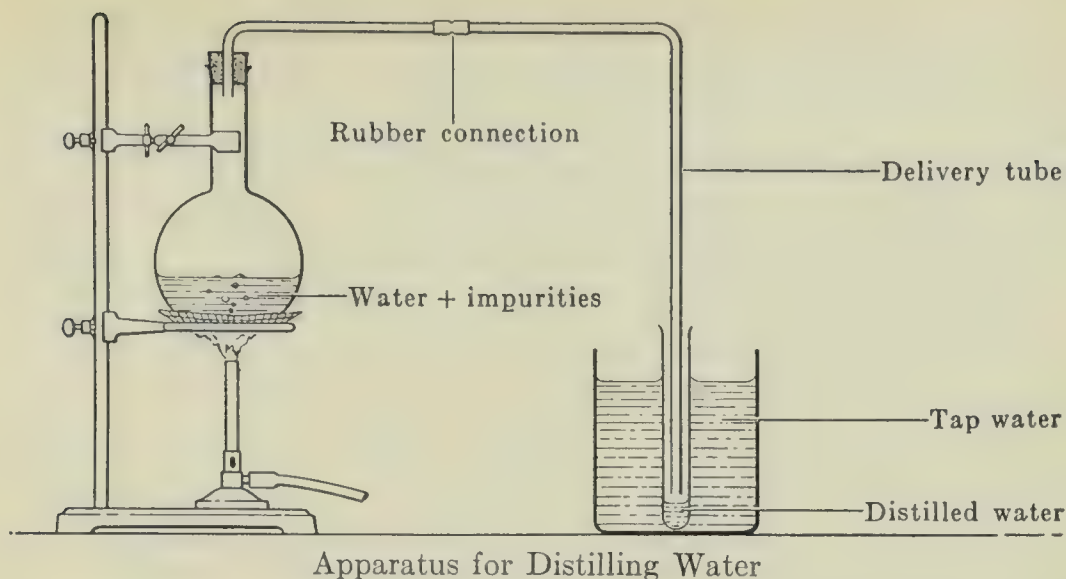
Apparatus for Distilling Water Using a Liebig Condenser

NOTE: While this experiment is in progress, or as a supplement to it, it is suggested that the distillation (with a Liebig Condenser) of some brightly colored water be carried on in clear view of the entire class.

1. Arrange apparatus as shown in the figure on page 85. Place half a teaspoonful of copper sulfate and an equal quantity of table salt in the flask. Add water to fill the flask almost $\frac{1}{2}$ full. Shake until both solutes are dissolved. Heat gently at first and then boil until $\frac{1}{2}$ of a test-tubeful of distillate is collected. Is the distillate colored? Taste it. Does it have a salty taste?

2. To some pure water in a test tube add a drop of phenolphthalein solution. Any result? Add 1 drop of dilute ammonium hydroxide (solution of ammonia gas in water). Result? This color change is a very sensitive test for the presence of ammonia water or other bases.

Empty the flask used in Exp. 1 above and wash it thoroughly.



Add water to fill it about $\frac{1}{3}$ full. Then add 2 or 3 drops of dilute ammonium hydroxide. Use the same apparatus as in Exp. 1. Heat the contents of the flask to the boiling point. Collect $\frac{1}{3}$ of a test-tubeful of distillate in each of 4 test tubes arranged in order of collection. Add just one drop of phenolphthalein solution to each test tube. Note any differences in shade of color resulting. If the last test tube still shows presence of ammonia, collect several $\frac{1}{2}$ test-tubefuls of distillate until one is obtained which shows the absence of ammonia.

Copy the following in your notebook, supplying the word or words needed to make each a complete statement. Underline the words you supply.

Observations and Questions on Experiments

1. In Exp. 1 the distillate is (colored, colorless) —?—, thus showing the (presence, absence) —?— of copper sulfate.
2. The distillate in Exp. 1 (does, does not) —?— taste salty, thus showing the (presence, absence) —?— of salt.
3. The liquid in the distillate in Exp. 1 is called —?— water and is formed by the —?— of the steam produced from boiling water. Steam is water in the —?— state.
4. Copper sulfate and salt are (more volatile than, less volatile than, as volatile as) —?— water.

5. The process of distillation (is effective, is not effective) —?— in ridding water of impurities such as copper sulfate and salt.

6. Phenolphthalein is colored —?— in the presence of ammonia water or any other base.

7. When a solution of water and ammonia gas is boiled, the —?— vaporizes first.

8. Ammonia is (more volatile than, less volatile than, as volatile as) —?— water.

9. The later portions of distillate from a solution of water and ammonia gas contain (more, less, the same amount of) —?— ammonia than the first portion. This is shown by —?—.

10. To rid water of ammonia gas, the method of distillation (is effective, is not effective) —?—.

Conclusions

1. In distillation a liquid is converted by —?— to the —?— state and then converted back again by —?— to the liquid state.

2. By distillation we can effectively rid water of impurities that —?—.

3. It is not good practice to use the process of distillation to rid water of gaseous or volatile impurities. Water can be rid of such impurities by —?—.

4. It is not necessary to use the process of distillation to rid water of undissolved impurities. Water can be rid of such impurities by —?—.

Supplementary Exercises

1. Pure water boils at 100° C. and freezes at 0° C. under standard conditions. If the water has dissolved impurities in it, the boiling point is (100° C., higher than 100° C., lower than 100° C.) —?— and the freezing point is (0° C., higher than 0° C., lower than 0° C.) —?—.

2. Alcohol boils at 78° C. If a mixture of equal amounts of water and alcohol is heated in the distilling flask, the liquid will start boiling when the temperature is (more than, less than, just) —?— 100° C. The first portion of the distillate will contain

mostly —?— and the last portion will be mostly —?—. This process of separating the ingredients of a mixture, which differ in their boiling points, into parts, by distilling the mixture and collecting the distillate into these separate parts is called —?—. The most important commercial use of this process is —?—.

3. Distilled water, rather than water from the faucet, is used in chemical laboratories for making solutions and for analytical purposes because —?—.

4. Rain water (is similar to, is not similar to) —?— distilled water because —?—.

5. A sample of water can be shown to have dissolved impurities by —?—.

6. Wholesome water (is, is not) —?— always chemically pure water because —?—.

7. Modern ocean liners obtain drinking water from sea water by —?—.

8. A glass of cold or iced water becomes misty on the outside because —?—.

9. Copy both List A and List B in your notebook. Write alongside each item in List A the number of the item in List B which corresponds. Do not write any number more than once.

List A

List B

- | | |
|---------------------------|---|
| (a) Filtration | 1. softens temporary hard water. |
| (b) Coagulation | 2. rids water of gases and kills bacteria. |
| (c) Aeration | 3. rids water of undissolved impurities. |
| (d) Distillation | 4. softens permanent hard water. |
| (e) Chlorination | 5. to make vichy or seltzer. |
| (f) Heating | 6. rids water of nonvolatile impurities. |
| (g) Adding limewater | 7. rids water of colloidal impurities. |
| (h) Adding carbon dioxide | 8. kills bacteria in water. |
| (i) Adding washing soda | 9. to aid thyroid gland deficiency. |
| (j) Adding sodium iodide | 10. kills bacteria and improves taste of water. |

Optional Questions

1. Why is water used in the cooling system of an automobile?
2. What is the boiling point of pure water? the freezing point? the temperature of greatest density?
3. Is clear water always fit to drink? Explain. Is water containing impurities always unfit for drinking? Explain.
4. What is sublimation? Give an illustration.
5. What is a precipitate? a filtrate? a distillate?
6. Why is water aerated?
7. How are bacteria in water destroyed?
8. Write an imaginary story of the travels of a molecule of water. Start with the molecule of water in the ocean being evaporated and blown in toward land.
9. How would you obtain a solid from its water solution? a gas from its water solution?
10. Prepare a report on the methods used in purifying the water of your home city.

UNIT 22. HYDROLYSIS

Experiments

NOTE: *This experiment will show that many salts are not neutral, that is, their solutions affect litmus paper.*

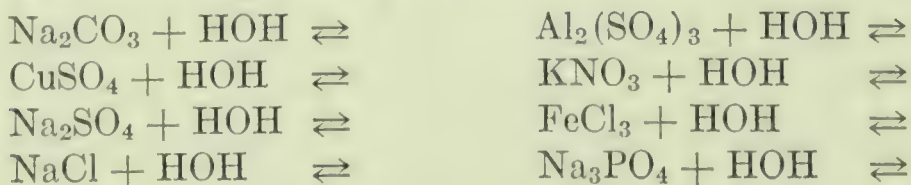
Place a small quantity of each of the solutions of the salts listed in the table below in separate clean test tubes. With a clean stirring rod, place a drop of each of the solutions on both red and blue litmus paper. (Be sure to rinse the stirring rod for each new salt solution.) Make a table in your notebook like the one below. Record your results in it.

WATER SOLUTION OF	FOR- MULA OF SALT	EF- FECT ON LITMUS	ACID, BASIC, OR NEU- TRAL REACTION	FOR- MULA OF ACID DERIVED FROM	FOR- MULA OF BASE DERIVED FROM	CON- CENTRA- TION OF HYDRO- GEN IONS (CH^+) (HIGH, LOW)	CON- CENTRA- TION OF HY- DROXYL IONS (COH^-) (HIGH, LOW)	WHICH DOM- INATES: H^+ or OH^- or NEITHER
✓ Sodium carbonate	Na_2CO_3	red to blue	basic	H_2CO_3	NaOH	low CH^+	high COH^-	OH^-
✓ Copper sulfate								
✓ Sodium sulfate								
✓ Sodium chloride								
✓ Aluminum sulfate								
Potassium nitrate								
Ferric chloride								
✓ Sodium phosphate								
Zinc sulfate								
Sodium tetraborate (borax)								

Copy the following in your notebook, supplying the word or words needed to make each a complete statement. Underline the words you supply.

Observations and Questions on Experiments

1. The —?— ions cause an acid reaction.
2. The —?— ions cause a basic or alkaline reaction.
3. A strong acid is one which ionizes —?— when in water solution and produces a (low, high) —?— concentration of —?— ions. The three strong acids that we study in elementary chemistry are: —?—. Most of the other acids can be considered as weak.
4. A strong base is one which ionizes —?— when in water solution and produces a (low, high) —?— concentration of —?— ions. The two strong bases that we study in elementary chemistry are —?—. Most of the other bases can be considered as weak.
5. Write the ionic equation which expresses the slight ionization of water.
6. This slight ionization of water makes possible reversible reactions with many salts to produce a —?— and an —?—.
7. All hydrolysis reactions are reversible. Copy and complete the following hydrolysis equations. Encircle the strong acids and the strong bases produced.



Conclusions

1. Hydrolysis is the reversal of a —?— reaction and is due to the slight —?— of water.
2. The water solution of a salt derived from a strong base and a weak acid gives (an acid, a basic, a neutral) —?— reaction. An example of such a salt is —?—.
3. The water solution of a salt derived from a weak base and a strong acid gives (an acid, a basic, a neutral) —?— reaction. An example of such a salt is —?—.

4. The water solution of a salt derived from a strong base and a strong acid gives (an acid, a basic, a neutral) —?— reaction. An example of such a salt is —?—.

NOTE: *In the case of a salt derived from a weak base and a weak acid, the base and acid are generally not equally weak and one usually dominates.*

Supplementary Exercises

1. A water solution of soap changes litmus from red to blue. Soap is not a base. It therefore is a salt of a (strong, weak) —?— acid and a (strong, weak) —?— base.

2. Ammonium chloride (NH_4Cl) solution changes litmus from blue to red. Ammonium chloride is a salt of the (strong, weak) —?— base —?— and of the (strong, weak) —?— acid —?—.

3. Indicate whether each of the following salts will give an acid, a basic, or a neutral reaction in water solution: (a) $\text{Zn}(\text{NO}_3)_2$, (b) K_2S , (c) $\text{NaC}_2\text{H}_3\text{O}_2$, (d) K_2SO_4 , (e) NaNO_3 , (f) MgCl_2 .

4. From the point of view of composition salts may be classified as normal salts, acid salts, or basic salts. An example of a normal salt is —?—. Its formula is —?—.

5. An example of an acid salt is —?—. Its formula is —?—.

6. An example of a basic salt is —?—. Its formula is —?—.

7. Normal salts are simple combinations of a —?— element or group and a —?— element or group.

8. The formula of an acid salt contains the element —?—.

9. The formula of a basic salt contains the —?— group.

10. An acid salt usually has (an acid, a basic, a neutral) —?— reaction when dissolved in water. An exception is the acid salt —?— which has a mild basic reaction in water solution.

11. A basic salt usually has (an acid, a basic, a neutral) —?— reaction when dissolved in water.

12. A normal salt as shown in the hydrolysis experiments may have either a —?— or a —?— or an —?— reaction when dissolved in water.

Optional Questions

1. Why does sodium carbonate make a good cleaning agent?
2. Explain why a water solution of sodium bicarbonate gives an alkaline reaction.
3. Describe the Solvay process for making sodium bicarbonate and sodium carbonate. Mention uses for each of these products.
4. What happens to starch when it undergoes the process of hydrolysis? Is the reaction similar to the hydrolysis of sodium carbonate? Explain.

UNIT 23. WHY REACTIONS GO TO COMPLETION

Experiments

1. Formation of a precipitate.

(a) To a small amount of calcium chloride solution (5 cc.) in a test tube add some silver nitrate solution. Result?

(b) To a small amount of sodium sulfate solution (5 cc.) in a test tube add some barium chloride solution. Result?

2. Formation of a gaseous product.

(a) To a little solid sodium chloride in a test tube add a little sulfuric acid (2:1). Warm gently. Test escaping vapors with moist blue litmus. Result?

(b) To a little solid ammonium chloride in a test tube add some sodium hydroxide solution. Warm gently. Smell cautiously. The gas evolved is ammonia gas. Test escaping vapors with moist red litmus. Result?

3. Formation of a nonionized or slightly ionized product.

Pour about $\frac{1}{2}$ a test-tubeful of dilute sodium hydroxide into a beaker. Neutralize it with a dilute solution of hydrochloric acid. (See procedure of Unit 18, Exp. 3.) Feel the outside of the beaker. Any change of temperature? Taste the mixture. Any evidence of reaction?

4. Reversible reactions.

(a) To a small amount of sodium nitrate solution (5 cc.) in a test tube add some barium chloride solution. Any evidence of a reaction?

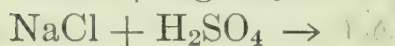
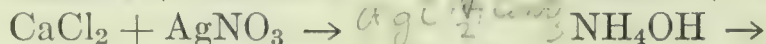
(b) To a small amount of sodium chloride solution (5 cc.) in a test tube add some potassium nitrate solution. Any evidence of reaction?

Copy the following in your notebook, supplying the word or words needed to make each a complete statement. Underline the words you supply.

Observations and Questions on Experiments

1. The observable evidence of chemical reactions in Exp. 1 is the formation of AgCl precipitate.

2. Copy and complete the following equations. Indicate a precipitate by an arrow pointing downward (\downarrow) and a gas by an arrow pointing upward (\uparrow).



3. The reactions in Exp. 2 are not reversible because they each produce a —?— product which leaves the field of action.

4. The observable evidence of the neutralization reaction (in Exp. 3) is the evolution of —?— and the —?— taste of the mixture.

5. Copy and complete the following equation and then write it in ionic form.



6. The —?— formed in Ex. 5 above ionizes to a negligible degree and thus in effect leaves the field of action.

7. Copy and complete the following reversible reactions and then write them in ionic form.



8. In reversible reactions there seems to be (evidence, no evidence) —?— of chemical reaction. The condition is called one of dynamic equilibrium because —?—.

Conclusions

1. Most chemical reactions are reversible. In reversible reactions the products —?— to reform the —?—. Both actions continue simultaneously. A state of —?— exists.

2. Some chemical reactions go to an end or completion because one of the products leaves the field of action as a —?—, or as a —?—, or as a —?— substance like water.

Supplementary Exercises

1. The concentration of a substance is the amount of this substance per unit of volume (e. g., —?— per —?—).

2. The equilibrium point of a reversible reaction can be shifted by increasing the concentration of one of the reacting substances, (Law of mass action). In the reaction:



if the concentration of NaCl is increased (by the addition of —?—), then the reaction will proceed toward —?—. If the concentration of Cl^- were increased (by addition of HCl), then the reaction would proceed toward —?—.

3. The equilibrium point of a reversible reaction can also be shifted by changing the temperature. For increase of temperature the reaction will tend to proceed in the direction which (absorbs, evolves) —?— heat and for decrease in temperature the reaction will tend to proceed in the direction which (absorbs, evolves) —?— heat. (Van't Hoff's principle.)

4. The following reactions indicate six ways of making sodium chloride. Copy and complete the table.

EQUATION	TYPE OF REACTION	REASON FOR REACTION GOING TO COMPLETION
..... + \rightarrow NaCl	Direct combination	The NaCl cannot decompose under ordinary conditions.
$\text{NaClO}_3 \rightarrow$ +		A gas is formed which leaves the field of action.
$\text{Na} + \text{HCl} \rightarrow$ +		
$\text{NaOH} + \text{HCl} \rightarrow$ +		
$\text{Na}_2\text{SO}_4 + \text{.....} \rightarrow \text{BaSO}_4 + \text{.....}$		
$\text{Na}_2\text{CO}_3 + \text{.....} \rightarrow \text{.....} + \text{H}_2\text{CO}_3$ $\text{H}_2\text{CO}_3 \rightarrow \text{H}_2\text{O} + \text{CO}_2$		

5. The equilibrium point of a reversible reaction can also be shifted by changing the pressure. This is only true for (gases, liquids, solids) —?—. For increase of pressure that reaction which causes (an increase, a decrease) —?— of pressure will be

avored. (Le Chatelier's principle—a generalization which also embraces Van't Hoff's principle.)

6. All neutralization reactions in effect go to completion because —?—.

NOTE: *The following solubility and insolubility rules are generally true for the common inorganic compounds and are useful in predictions concerning reversibility.*

7. Most chlorides are (soluble, insoluble) —?— in water. Three exceptions are: —?—.

8. Most sulfates are (soluble, insoluble) —?— in water. Three exceptions are —?—.

9. Practically all —?—, acetates, and chlorates are soluble in water.

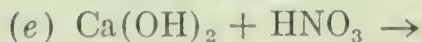
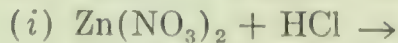
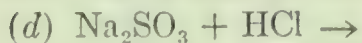
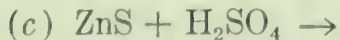
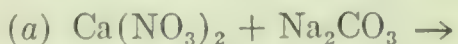
10. Most carbonates are (soluble, insoluble) —?— in water. Three exceptions are —?—.

11. Most sulfides and hydroxides are (soluble, insoluble) —?— in water. Some exceptions are the compounds of the five metals: —?—.

12. Most metallic oxides are (soluble, insoluble) —?— in water. Four exceptions are: —?—. The soluble metallic oxides react with water to form soluble —?—.

Optional Questions

1. Copy and complete the following equations. Which reactions go to completion? Indicate the reason in each case. Indicate the reversible reactions by a double arrow.



2. Predict whether a reaction will take place between each of the following pairs of substances. Give a reason for your prediction. If you think a reaction would occur upon application of heat, electricity, or some other form of energy, or under special conditions, so indicate. Write equations for those you think will react.

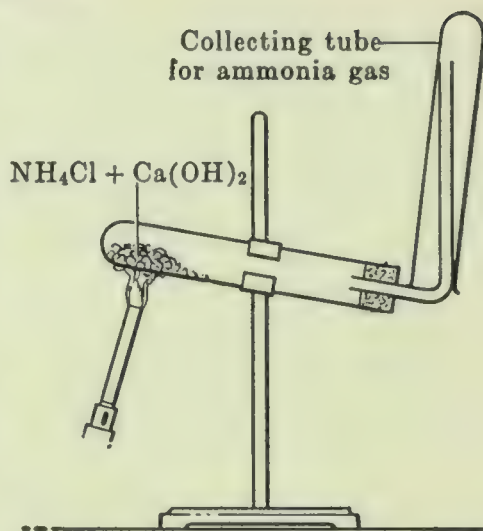
- | | |
|---------------------------------|------------------------------------|
| (a) copper + zinc sulfate | (g) gold + nitrogen |
| (b) magnesium + sulfur | (h) sodium chlorate (heated) |
| (c) tin oxide (heated) | (i) sodium nitrate + sulfuric acid |
| (d) sodium + aluminum chloride | (j) potassium chloride + sodium |
| (e) hydrogen sulfide + chlorine | nitrate |
| (f) copper + hydrogen | |

3. State and illustrate the law of mass action.

UNIT 24. AMMONIA

Experiments

1. Place about one-quarter of a test-tubeful of solid ammonium chloride (sal ammoniac) (9 g.) on a clean piece of paper. Odor? On another piece of paper place the same bulk of dry slaked lime (calcium hydroxide) (7 g.). Mix together. Rub a little of this mixture between your fingers. Does the mixture have an odor? Hold a piece of moist red litmus paper over some of this rubbed mixture. Result? Place the mixture in a dry test tube and arrange the apparatus as shown in the figure at the right. Heat gently.



2. Collect a test-tubeful of gas. Use a dry test tube. Cover the mouth of the test tube with your thumb and then invert in a dish of water. Remove thumb. Result?

3. Collect a test-tubeful of gas. Test with a burning splint. Result?

4. Collect a bottle of gas. Pour a test-tubeful of water into the bottle. Place the palm of your hand over the bottle and shake. The resulting solution is a solution of ammonium hydroxide. Pour some of this solution into a test tube and heat to boiling. Test the escaping vapors with wet red litmus paper. Result?

5. Pour a drop or two of concentrated hydrochloric acid into a test tube so as to wet the sides. Invert and place over the delivery tube (while gently heating the generating test tube). Result?

6. To a small amount of ammonium sulfate in a test tube add a few drops of sodium hydroxide solution. Heat gently. Note the odor. All ammonium compounds behave similarly.

Copy the following in your notebook, supplying the word or words needed to make each a complete statement. Underline the words you supply.

Observations and Questions on Experiments

1. The ammonium chloride (has, has not) —?— an odor. The slaked lime (has, has not) —?— an odor. The mixture of ammonium chloride and slaked lime has an odor of —?—.

2. Moistened red litmus held over a mixture of ammonium chloride and calcium hydroxide turns —?—, indicating the presence of a —?—.

3. Ammonia is collected by the method of —?— because it is ~~lighter~~ —?— and —?—.

4. In the preparation of ammonia the reaction takes place in two stages. Copy and complete the following equations:



5. When a test-tubeful of ammonia gas is inverted in a dish of water, the result is ~~the gas dissolves~~. This shows that ammonia is —?— in water.

6. Ammonia (is, is not) —?— combustible and (does, does not) —?— support combustion.

7. The equation for the reaction of ammonia and water is: —?—.

8. The reaction in Ex. 7 above will proceed toward the right if the temperature is —?— and will proceed toward the left if the temperature is —?—.

9. Ammonium hydroxide is (a stable, an unstable) —?— compound.

10. Concentrated hydrochloric acid is volatile and evolves ~~hydrogen chloride~~ gas. This gas reacts with ammonia gas to form a white smoke consisting of fine solid particles of ~~ammonium chloride~~. The equation for this reaction is: —?—.

11. The odor of —?— is noticeable when an ammonium salt is heated with a base.

12. The equation for the reaction between ammonium sulfate and sodium hydroxide is: —?—.

Conclusions

1. Ammonia is prepared in the laboratory by the reaction between an NH_4Cl salt and a Ca(OH)_2 .

2. The equation for the laboratory preparation of ammonia is: $\text{NH}_4\text{Cl} + \text{Ca(OH)}_2 \rightarrow \text{—?—} + \text{CaCl}_2 + 2\text{H}_2\text{O}$

3. An outstanding chemical property of ammonia is its reaction with H^+ to form a volatile base called NH_4^+ .

4. To test for an ammonium compound, heat some of the unknown substance with a solution of —?— . If the odor of —?— is given off, the unknown was an ammonium compound. The gas evolved will also change wet litmus from —?— to —?— .

Supplementary Exercises

1. A good deal of the ammonia of industry is obtained as a byproduct in the manufacture of CaCl_2 from —?— .

2. Ammonia is also made commercially by the direct synthesis of —?— and —?— under special conditions of temperature and pressure and in the presence of a catalyst such as iron. This is known as the —?— process. The equation is: —?— . Reversal of this reaction is lessened by dissolving the ammonia formed in —?— .

3. In the cyanamid process, ammonia is made by passing nitrogen over hot —?— , forming calcium cyanamid. This is then treated under pressure with —?— to form ammonia.

4. The fixation of atmospheric nitrogen means —?— .

5. Ammonia is used as a refrigerant in artificial ice plants. First the gas is converted into —?— by means of —?— . The ammonia is then allowed to —?— thus producing the effect of —?— .

6. Ammonia is used as a cleanser by dissolving it in H_2O to form —?— . This cleans by —?— grease and thereby loosening dirt. One important advantage it has over most other basic cleaners is —?— .

7. Ammonia is also used in the Solvay process for manufacturing —?— .

8. Ammonia is often evolved in the decay of —?— .

9. The ammonium radical differs from the gas ammonia in that the former has the symbol NH_4^+ and (does, does not) —?— exist as a compound while the latter has the formula NH_3 and (does, does not) —?— exist as a compound.

10. Household ammonia should be kept well-stoppered and in a cool place because NH_3 is very soluble in water.

11. Ammonium hydroxide is preferable to sodium hydroxide in neutralizing an acid spot on clothing because it —?—.

12. The presence of ammonium compounds in drinking water is a sign of decay because NH_4^+ is produced by the decay of organic matter.

Optional Questions

1. How does the ammonium radical differ from ammonia? From a metal?

2. Explain the use of ammonia in the artificial manufacture of ice. Why is brine used in the process?

3. Explain the use of ammonia in the gas type of refrigerator.

4. Why would you not dry ammonia gas by bubbling it through concentrated sulfuric acid?

5. What type of organic matter will form ammonia in the process of decay?

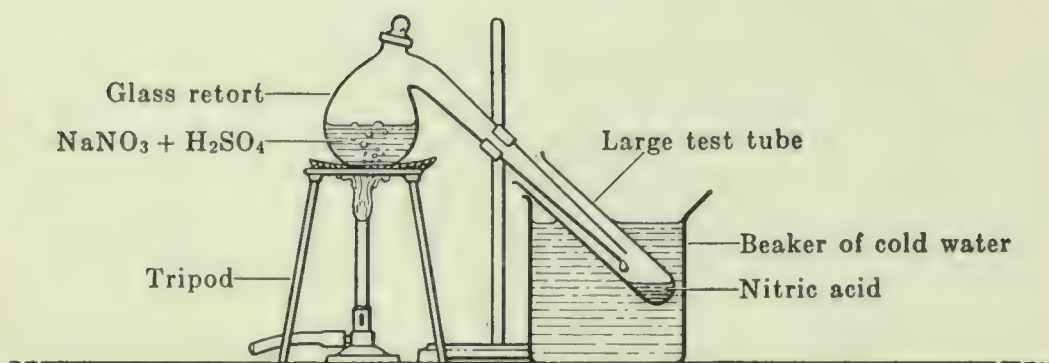
6. What are smelling salts?

7. Why is ammonium hydroxide called the volatile alkali?

8. How can a knowledge of the law of mass action be applied to make the preparation of ammonia by the Haber process more efficient and more economical?

UNIT 25. NITRIC ACID

Experiments



Apparatus for Preparing Nitric Acid

Caution: Concentrated sulfuric acid and nitric acid are dangerous. Do not get them on your skin or clothing.

1. Place about a teaspoonful (9 g.) of sodium nitrate (Chile saltpeter) in a retort. Arrange apparatus as shown in the figure above. Through a funnel, pour $\frac{1}{2}$ a test-tubeful of concentrated sulfuric acid upon the sodium nitrate. Stopper the retort and heat gently. (If overheated some of the nitric acid produced will decompose to form brown nitrogen dioxide gas which discolors the nitric acid.) Collect about 6 cc. of the nitric acid. (When the retort has cooled, add warm water to dissolve the residue.)

2. Add a drop of the nitric acid to 10 cc. of water in a test tube. Test with litmus. Result?

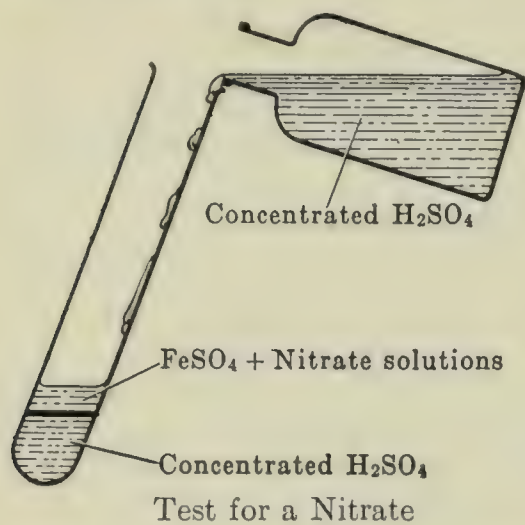
3. Pour 2 cc. of the nitric acid into a test tube and heat cautiously. While the tube is being heated, hold a glowing wood splint above the acid. Result?

4. Place a small piece of the white of a cooked egg in a test tube. Put a drop of the nitric acid on it. What color effect? Nitric acid has this effect on all protein matter.

5. Pour 1 cc. of concentrated nitric acid on a small piece of copper in a test tube. Result?

6. Make dilute nitric acid by adding 1 cc. of the concentrated nitric acid you prepared in Exp. 1 to 3 cc. of water in a

test tube. Drop a small piece of copper into the dilute nitric acid. Heat gently. Result?



7. Dissolve a pinch of sodium nitrate in 4 cc. of water. Add 4 cc. of freshly prepared ferrous sulfate solution. Hold the test tube with a test tube holder in an inclined position and carefully pour down the side of the test tube 4 cc. of concentrated sulfuric acid. The heavy sulfuric acid will form a separate layer on the bottom. What

appears between the two layers? All nitrates give this effect.

Copy the following in your notebook, supplying the word or words needed to make each a complete statement. Underline the words you supply.

Observations and Questions on Experiments

1. A glass retort is used in this experiment because —?—.
2. Nitric acid is evolved in the —?— state and is condensed to its normal —?— state.
3. Sulfuric acid is used to supply the —?— and because it has a —?— boiling point. Sodium nitrate is used to supply the —?—. Both are comparatively inexpensive.
4. The equation for the reaction is: $NaNO_3 + H_2SO_4 \rightarrow$ —?—.
5. The reaction in Ex. 4 is not reversible because —?—.
6. Nitric acid changes litmus from —?— to —?—, indicating the presence of —?— ions.
7. Nitric acid is a (weak, strong) —?— acid, because it produces —?— ions in water solution.
8. The fumes from boiling nitric acid cause the glowing wood to —?—. This shows that boiling nitric acid gives off —?—.

9. Complete the equation: HNO_3 (concentrated) $\xrightarrow{\text{heated}}$
10. Pure nitric acid is colorless. The nitric acid prepared in the laboratory is colored yellow because —?—.
11. Concentrated nitric acid colors white egg albumen —?—. Woolen clothing and the skin are similarly affected. Nitric acid stains all —?— compounds this way.
12. When concentrated nitric acid reacts with copper, the bluish-green substance produced is —?—. The brown gas produced is —?—.
13. Copper (can, cannot) —?— replace hydrogen from acids because it is (more active, less active) —?— than hydrogen.
14. Copper reacts with nitric acid because the nitric acid is —?— agent.
15. Complete the equations:
 $\text{Cu} + \text{HNO}_3(\text{concentrated}) \rightarrow$ $\text{Cu} + \text{HNO}_3(\text{dilute}) \rightarrow$
16. In testing for a nitrate, a —?— appears between the bottom layer of —?— and the top layer of the mixture of —?— and —?—.

Conclusions

1. Nitric acid is prepared in the laboratory (and also commercially) by the reaction between —?— and —?—. The nitric acid is evolved in the —?— state and then —?— to the normal —?— state.
2. Nitric acid is not only a —?— but also has the additional chemical property of being a powerful —?— agent.
3. When dilute nitric acid reacts with metals, the three products formed are: —?—. When concentrated nitric acid reacts with metals, the three products formed are: —?—.
4. In testing for a nitrate, add some freshly prepared —?— to some of the unknown substance in solution. Then carefully pour down the side of the test tube some concentrated —?—. A

—?— in between the two layers indicates the presence of a nitrate.

Supplementary Exercises

1. A mixture of concentrated nitric and hydrochloric acids is called —?—. Platinum and gold, which are not affected by single acids, dissolve in this mixture because the —?— produced when the two acids react forms soluble —?— with these metals.
2. Ordinarily when a metal reacts with an acid, —?— and —?— are formed. —?— is not liberated when nitric acid reacts with a metal because —?—.
3. Four important uses of nitric acid are —?—.
4. In the arc (Birkeland-Eyde) process for the commercial manufacture of nitric acid, the —?— and —?— of air are combined with the aid of an arc discharge of electricity. The resulting —?— reacts with more —?— from the air and with —?— to form nitric acid. Write the two equations.
5. In another commercial process for making nitric acid, ammonia made synthetically by the —?— process is oxidized in the presence of a catalyst (platinum). The resulting —?— is caused to react with —?— of air and with —?— to produce nitric acid. Write the two equations.
6. The fixation of atmospheric nitrogen means the —?—.
7. A government plant for the fixation of atmospheric nitrogen was located at Muscle Shoals because —?—.
8. Nitrogen compounds are used in explosives because —?—.
9. An important nitrogen compound found as a mineral deposit is —?—. It is commonly known as —?—.
10. There was an enormous demand for nitric acid during the World War because —?—.
11. The development of artificial methods of making nitrates is necessary because —?—.
12. Copy and complete the table of the oxides of nitrogen at the top of the next page.

FORMULA	CHEMICAL NAME	COLOR	ACID ANHYDRIDE OF	EQUATION FOR PREPARATION
N_2O			×	
NO or N_2O_2			×	
NO_2 or N_2O_4				
N_2O_3				×
N_2O_5				×

13. Nitrous oxide, commonly known as —?—, is used as —?—.

14. The oxides of nitrogen listed above form a series of compounds which illustrate very well the law of —?—.

Optional Questions

1. Explain how nitrates are formed in nature. Why are nitrates not found abundantly as mineral deposits?

2. How could you distinguish chemically between nitrous oxide and oxygen? Between nitrous oxide and nitric oxide?

3. What commercial method for making nitric acid is used in Norway? Why?

4. Why was there such demand for nitrates during the World War? Why were artificial fixation of nitrogen processes developed by the warring powers?

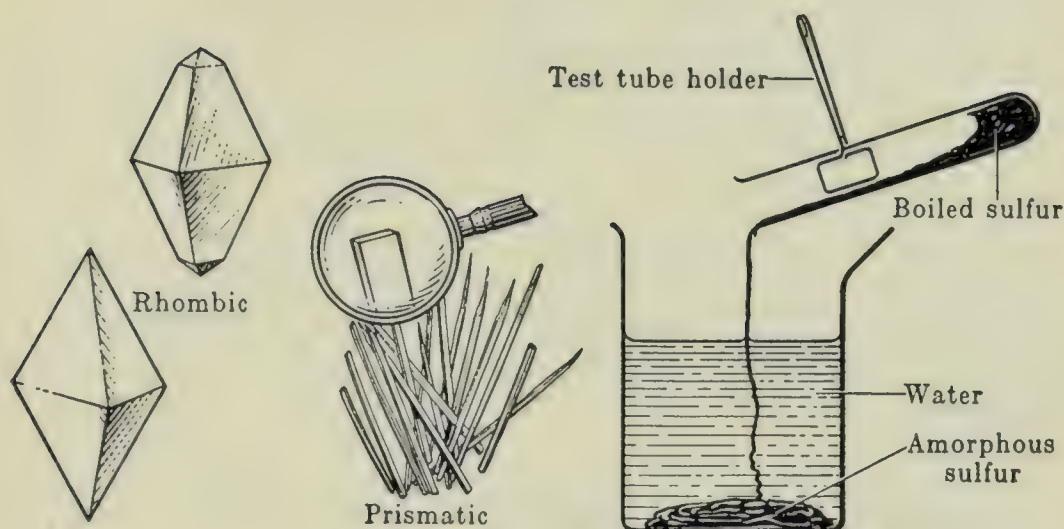
5. Name some explosives made with the use of nitric acid. Tell how two of these are made. Why do modern explosives contain nitrogen compounds?

6. Give the formula of nitrous acid. Give the names and formulas of three of its salts.

7. Write a 300-word essay on the various methods of nitrogen fixation employed in this country.

UNIT 26. SULFUR

Experiments



The Allotropic Forms of Sulfur

Caution: Carbon disulfide is very inflammable. Keep away from flames.

1. Allotropic forms of sulfur:

(a) To a test tube containing a pinch of powdered roll sulfur add 2 cc. of carbon disulfide. Shake until the sulfur dissolves. Pour the clear liquid on a watch glass; cover with a sheet of paper; and set aside (**away from flames**) to evaporate. (In the meantime, proceed with the next experiment.) Examine the residue with a magnifying lens. It consists of crystals of rhombic sulfur. Examine again after a few days.

(b) Gently heat $\frac{1}{2}$ of a test-tubeful (15 g.) of powdered sulfur until it melts. Be sure that the sulfur does not become darker than a pale yellow. Have ready a piece of filter paper folded so that it could fit into a funnel. Pour the melted sulfur into the filter paper. When a crust forms on the surface, spread out the filter paper. The crystals that have formed are crystals of prismatic or monoclinic sulfur. Examine them with a magnifying lens. Examine again after a few days.

(c) Heat $\frac{1}{3}$ of a test-tubeful (10 g.) of sulfur until the sulfur boils. Pour out the melted sulfur into a dish of cold water. Examine the solidified sulfur with a magnifying lens. It is amorphous sulfur. Examine again after a few days.

2. Chemical properties of sulfur:

(a) Heat a small quantity of sulfur in a test tube until some of it vaporizes. While heating, drop in a thin strip of copper foil or wire. Result?

(b) Rub a little powdered sulfur on a silver coin. Result?

(c) *Demonstration*—Mix a **little** powdered sulfur with an equal bulk of powdered zinc in a little pile on an asbestos mat. **At arm's length** cautiously ignite the mound with a Bunsen flame. Result?

Copy the following in your notebook, supplying the word or words necessary to complete each statement. Underline the words you supply.

Observations and Questions on Experiments

1. Rhombic sulfur (has, has not) —?— a definite form or shape, is (an amorphous, a crystalline) —?— variety of sulfur, and can be made by —?— a solution of —?— in carbon disulfide.

2. The watch glass in which the sulfur dissolved in carbon disulfide is placed is covered with a sheet of paper in order to —?—.

3. Prismatic or monoclinic sulfur (has, has not) —?— a definite form or shape, is (an amorphous, a crystalline) —?— variety of sulfur, and can be made by slow cooling of —?—.

4. Amorphous or plastic sulfur (has, has not) —?— a definite form or shape and is made by —?— boiling sulfur suddenly.

5. The only form of sulfur that remains unchanged after standing at room temperature for a few days is the —?— form. The other two forms, namely, —?— and —?— sulfur, change to the —?— form upon standing at room temperature.

6. Sulfur behaves like a (metal, nonmetal) —?— in combining with copper. The reaction is (endothermic, exothermic) —?—. The equation is: $\text{Cu} + \text{S} \rightarrow \text{—?—}$.

7. Powdered sulfur rubbed on silver produces a —?— colored deposit of —?—. The equation is $\text{Ag} + \text{S} \rightarrow \text{—?—}$.

8. The combination of zinc and sulfur is accompanied by —?—.

9. Zinc combines (more, less) —?— readily with sulfur than does copper. The equation is: $\text{Zn} + \text{S} \rightarrow \text{—?—}$.

Conclusions

1. The three allotropic varieties of sulfur are: —?—. These forms differ mainly in their (physical, chemical) —?— properties.

2. Below 96°C . the stable form of sulfur is the —?— form. Between 96°C . and 114°C . (melting point of sulfur) the stable form is the —?— sulfur.

3. Sulfur is an active (metal, nonmetal) —?— and combines readily with (metals, nonmetals) —?— to form compounds called —?—.

Supplementary Exercises

1. Sulfur occurs in the free or uncombined state in volcanic regions such as —?— and —?—. The largest deposits of sulfur in the world occur in the United States in —?—.

2. The large United States deposits of sulfur occur underground and in many places underneath quicksand. In the —?— process for extracting sulfur in these localities a hole is drilled to the sulfur bed. Then reaching down to the sulfur are placed three concentric pipes, a 6-inch, a 3-inch, and a 1-inch pipe. Down the 6-inch pipe is sent —?— which —?— the sulfur. Down the 1-inch pipe is sent —?— which mixes with the —?— sulfur and produces —?—. Up the 3-inch pipe comes the —?— sulfur forced up by the pressure of both the —?— and the —?—.

3. Four important uses for sulfur are: —?—.

4. Allotropes of the same element differ mainly in —?— properties. Isotopes of the same element differ only in —?—.

5. An allotropic variety of oxygen is —?—.

6. When the allotropic varieties of sulfur are burned in air or oxygen, the product formed in each case is —?—. This indi-

cates a similarity in —?— properties. Complete the equation:
 $S + O_2 \rightarrow$

7. Flowers of sulfur is a commercial variety formed by —?—. It contains both —?— and —?— forms of sulfur.

8. Roll sulfur is a commercial variety made by —?—. It consists almost wholly of the —?— form of sulfur.

9. The tarnishing of silver in the home is usually caused by the action of —?— or its compounds in —?— or in the air with the formation of a —?— colored deposit of —?—.

10. Black gunpowder is a mixture of sulfur with —?— and —?—.

11. Oxygen (atomic No. 8) and sulfur (atomic No. 16) each contain —?— electrons in the outermost ring, indicating a —?— in their —?— properties. Each is a (metal, nonmetal) —?— with a valence of —?—. Each will (lend, borrow) —?— (1, 2, 3, 4) —?— electrons (to, from) —?— (metals, nonmetals) —?—.

Optional Questions

1. How does sulfur occur in nature? Mention some localities in which deposits are found.

2. Explain the reasons for using sulfur in the vulcanizing and making of rubber products. Who was the inventor of this process?

3. Explain the reasons for using sulfur in the making of matches.

4. Compare the energy content of the allotropic varieties of sulfur. How could you prove chemically that they contain sulfur and nothing else?

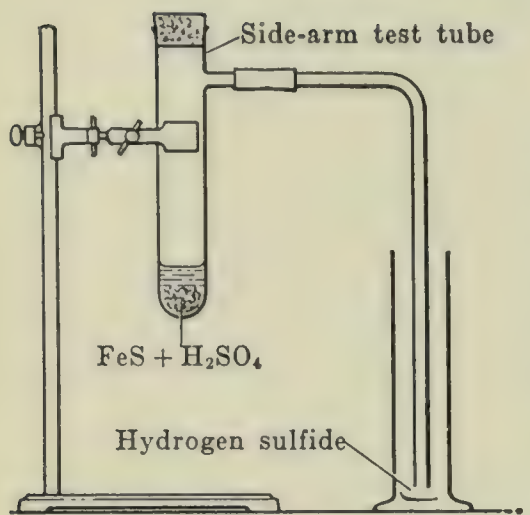
5. Mention 3 elements other than sulfur that exist in allotropic forms.

6. Explain the similarity of sulfur and oxygen from the electronic and chemical points of view.

7. Mention some foods that contain combined sulfur. What effect have these foods on silver utensils?

UNIT 27. HYDROGEN SULFIDE

Experiments



Apparatus for Preparing Hydrogen Sulfide

1. Arrange apparatus as shown in the figure at the left. Place a small quantity of ferrous sulfide in the generating test tube. Add dilute hydrochloric or sulfuric acid sufficient to cover the ferrous sulfide. Collect a dry tubeful of gas.

2. Bring this tube of gas to the Bunsen flame. What collects on the inside walls of the tube? The hydrogen sulfide is here being

burned in a limited supply of air (incomplete combustion).

3. Bubble some hydrogen sulfide through a test tube $\frac{2}{3}$ full of water for a few minutes. Test the solution with litmus paper. Taste it. Result?

4. Pour 5 cc. of the hydrogen sulfide solution made in Exp. 3 into 5 cc. of each of the following solutions set up in test tubes: copper sulfate, cadmium nitrate, antimony chloride, and zinc nitrate. Results? Colors?

5. Put a drop of hydrogen sulfide solution on a silver coin. Result?

6. Dip a piece of filter paper into a solution of lead acetate and hold it in the path of the hydrogen sulfide gas as it issues from the delivery tube. Result? This color change serves as a test for hydrogen sulfide gas.

7. Attach a jet tube or blowpipe to the delivery tube and, making sure there is a steady flow of gas (feel against your hand), ignite the gas. Cautiously smell the products of combustion. Hold a cold dry test tube over the flame. What condenses?

Copy the following in your notebook, supplying the word or words needed to make each a complete statement. Underline the words you supply.

Observations and Questions on Experiments

1. Hydrogen sulfide is collected by the method of —?— because it is —?— and —?—.

2. In making hydrogen sulfide, the acid is used to supply the —?— and the ferrous sulfide is used to supply the —?—. The equation for the reaction is: $\text{FeS} + \text{HCl} \rightarrow \text{—?—}$.

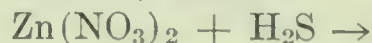
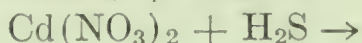
3. When hydrogen sulfide is burned in a limited supply of air, —?— collects on the sides of the test tube.

4. The equation for the incomplete combustion of hydrogen sulfide is: —?—.

5. A water solution of hydrogen sulfide changes litmus from —?— to —?— and has a —?— taste.

6. The test with litmus indicates the presence of —?— ions in a solution of hydrogen sulfide. It may therefore be classified as (an acid, a base, a salt) —?—.

7. Copy and complete the following equations. Indicate the name and color of each precipitate underneath its formula.



8. A silver coin —?— in contact with hydrogen sulfide. The compound formed is silver —?—.

9. Lead acetate paper turns —?— in the presence of hydrogen sulfide.

10. Hydrogen sulfide when burned with an abundant supply of air (complete combustion) yields —?— and —?—.

11. The equation for the complete combustion of hydrogen sulfide is —?—.

Conclusions

1. Hydrogen sulfide is prepared in the laboratory by the reaction between —?— and —?—.

2. Four physical properties of hydrogen sulfide are: —?—.

3. A water solution of hydrogen sulfide shows typical —?—

properties in that it has a —?— taste and that it changes litmus from —?— to —?—.

4. When —?— reacts with solutions of most metallic salts, there are formed insoluble —?— which precipitate out of solution. These precipitates can be differentiated by —?—.

Supplementary Exercises

1. We could substitute —?— for ferrous sulfide and —?— for hydrochloric acid in the preparation of hydrogen sulfide.

2. Another chemical name for hydrogen sulfide, recognizing its acid nature, is —?—.

3. Hydrogen sulfide is a (strong, weak) —?— acid because it produces —?— in water solution.

4. Hydrogen sulfide is useful to the analytical chemist because —?—.

5. Silver tarnishes at home when it reacts with the —?— compounds in foods. The tarnish that forms on the silver is —?—.

6. Lead paints when exposed to sulfur compounds turn —?— because of the formation of —?—.

7. Zinc paints retain their white color in the presence of sulfur compounds because —?—.

8. To test for a sulfide, —?— is added. The resultant evolution of —?— gas recognized by —?— and by its ability to turn lead acetate paper —?— shows the presence of a sulfide.

9. The blackening of silver in contact with a rubber article indicates that the rubber contains —?—.

Optional Questions

1. Describe the use of hydrogen sulfide in a chemical laboratory.

2. Hydrogen sulfide acts as a reducing agent in many chemical reactions. Explain.

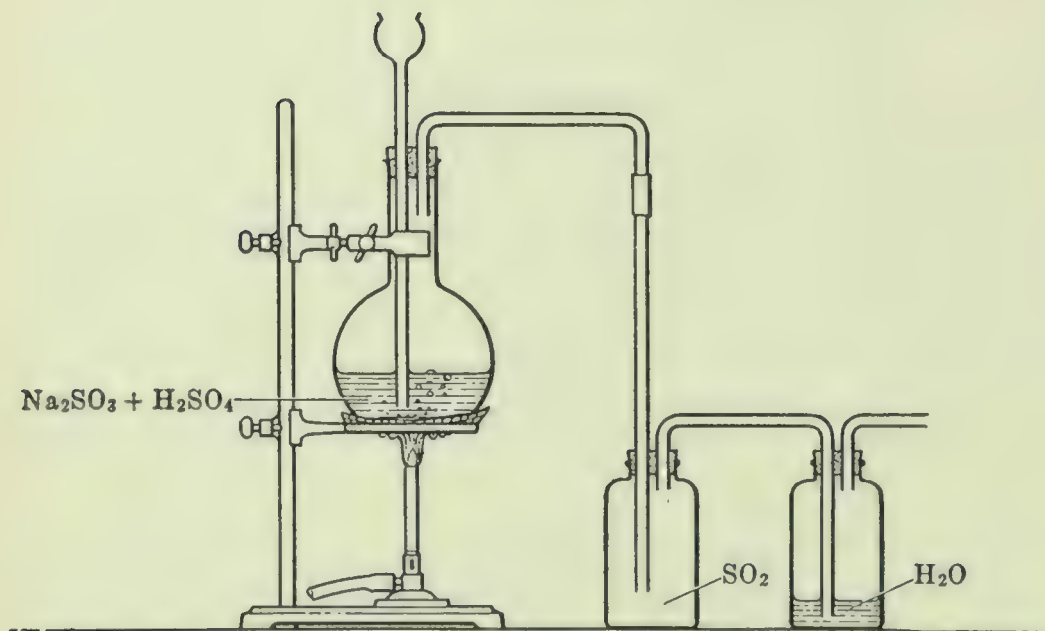
3. Explain the formation of sulfur particles in a solution of hydrogen sulfide allowed to stand for some time.

4. Why is hydrogen sulfide not prepared by the direct combination of hydrogen and sulfur?

5. With the aid of equations describe three ways of producing lead sulfide.

UNIT 28. SULFUR DIOXIDE

Experiments



Apparatus for Preparing Sulfur Dioxide

1. Preparation:

(a) Ignite a small piece of roll sulfur in a combustion spoon and lower into a wide mouth bottle. When the sulfur stops burning, cover the bottle with a glass plate. Note the color and **cautiously** note the odor of the sulfur dioxide gas produced. Insert a burning splint into the bottle. Result?

(b) Arrange apparatus as shown in the figure above. Place about 2 teaspoonfuls of sodium sulfite in the flask and pour down the thistle tube sufficient dilute sulfuric acid to cover the sodium sulfite. If necessary, heat gently. Collect 3 bottles of gas.

2. Properties:

(a) Pour a little water into a bottle of sulfur dioxide gas. Place the palm of your hand over the mouth of the bottle and shake. Test the liquid with blue litmus paper. Result? Taste it. Result?

(b) Pour a little dilute potassium permanganate solution into

another bottle of sulfur dioxide. Place the palm of your hand over the mouth of the bottle and shake. Result?

(c) Place a moist colored flower or a wet piece of red apple skin into a bottle of sulfur dioxide. Cover with a glass plate and allow to stand for a few minutes. Result?

Copy the following in your notebook, supplying the word or words needed to make each a complete statement. Underline the words you supply.

Observations and Questions on Experiments

1. Sulfur dioxide can be made by burning —?— in —?—. The equation is: —?—.

2. The reaction of Exp. 1 (b) above occurs in two stages. The equations are: —?—. This method is more convenient in the laboratory than the method of Exp. 1 (a) because —?—.

3. Sulfur dioxide is collected by the method of —?—.

4. Some of the physical properties of sulfur dioxide observed are: color, —?—; odor, —?—; solubility in water, —?—; density (compared with air) —?—.

5. Sulfur dioxide is (combustible, noncombustible) —?— and (does, does not) —?— support combustion.

6. The solution of sulfur dioxide in water changes litmus from —?— to —?— and has a —?— taste. This shows the presence of —?— ions in the solution.

7. Complete the equation: $\text{SO}_2 + \text{H}_2\text{O} \rightarrow$ —?—.

8. The sulfur dioxide in the reaction in Ex. 7 is called —?— anhydride.

9. A water solution of sulfur dioxide turns the color of potassium permanganate solution from —?— to —?—. The potassium permanganate reacts in this way when it is reduced (when its —?— is removed). The reducing agent here is —?— acid. When —?— acid takes on oxygen, it is converted into —?—.

10. Sulfur dioxide has the effect of —?— the colored flower (or red apple skin). Water or moisture is necessary in order for sulfur dioxide to react with it to form —?— which actually does the work of —?—.

Conclusions

1. Sulfur dioxide can be made commercially by burning —?— in —?—.

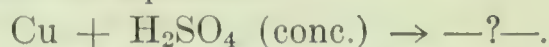
2. In the laboratory, sulfur dioxide is made by the reaction between —?— and —?—. In the resulting double replacement, —?— is formed which immediately decomposes into —?— and —?—.

3. An outstanding chemical property of sulfur dioxide is its ability to react with —?— to form the weak unstable —?— acid. Upon exposure to air or in the presence of an —?— agent, this —?— acid takes on —?— to form —?— acid.

4. Sulfur dioxide bleaches by first combining with —?— to form —?—. This substance actually does the bleaching by either removing —?— from the color or by combining with the color to form a colorless —?—.

Supplementary Exercises

1. Another method of preparing sulfur dioxide is by the reaction of hot concentrated sulfuric acid with a metal such as copper. Complete the equation:



The concentrated sulfuric acid here acts as —?— agent.

2. Both chlorine and sulfur dioxide need —?— in order to bleach. Sulfur dioxide rather than chlorine is used to bleach articles such as —?— because —?—.

3. Bleaching with sulfur dioxide is (more, less) —?— permanent than bleaching with chlorine.

4. Sulfur dioxide is used as the refrigerating agent in many electric refrigerators because —?—.

5. Two uses for sulfur dioxide other than those mentioned above are: —?—.

6. Sulfur or sulfur compounds should be removed from fuels such as coal gas because —?—.

7. In testing for a sulfite, add a few drops of —?— to the substance to be tested. The evolution of —?— gas which has a sharp, pungent odor and which will —?— the color of a dilute

potassium permanganate solution indicates the presence of a sulfite.

8. Sulfurous acid is a (weak, strong) —?— acid. This means that sulfurous acid does not —?— in water solution.

9. Sulfurous acid is (stable, unstable) —?— and is (volatile, nonvolatile) —?—.

10. Complete the equation: $\text{H}_2\text{O}_2 + \text{H}_2\text{SO}_3 \rightarrow \text{—?—}$. In this reaction the hydrogen peroxide is an —?— agent and the sulfurous acid is a —?— agent.

11. Newspapers and straw hats become “yellow with age” because —?—.

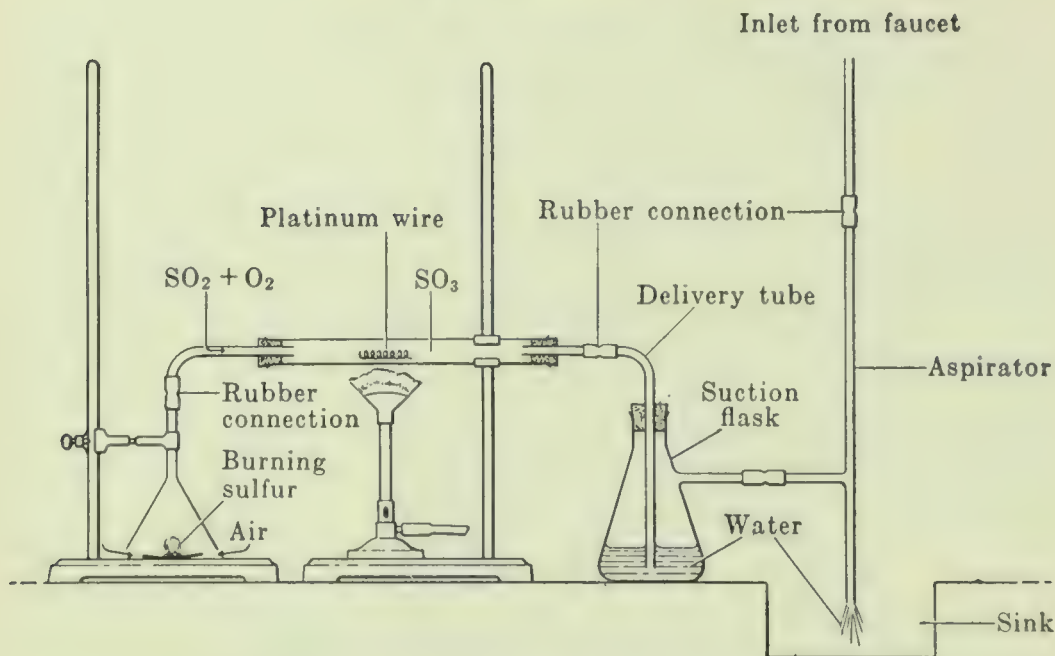
12. Sulfur being (metallic, nonmetallic) —?— is an (acid-forming, a base-forming) —?— element. The two equations for the preparation of H_2SO_3 from sulfur are: —?—.

Optional Questions

1. Give the names and formulas of 4 salts of sulfurous acid.
2. What is the formula of calcium bisulfite? How is it used in the manufacture of paper?
3. Give an illustration of sulfurous acid acting as a reducing agent.
4. Discuss the use of sulfur dioxide as a preservative of food.
5. Compare chlorine, sulfur dioxide, and hydrogen peroxide as bleaching agents.
6. Sulfur dioxide is stable, sulfurous acid is unstable. Explain.

UNIT 29. SULFURIC ACID

Experiments



Apparatus for Preparing Sulfuric Acid

1. Demonstration.

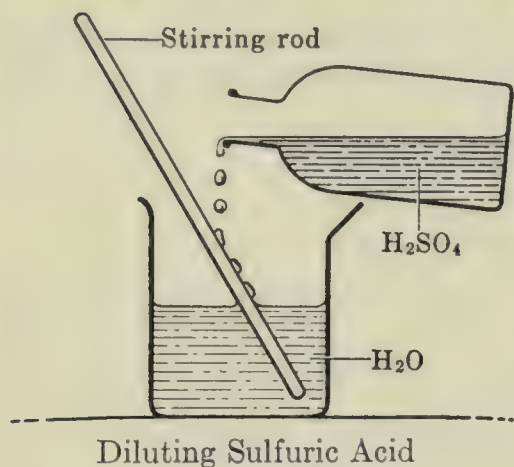
(a) Arrange apparatus as shown in the figure above. Turn on the water through the aspirator to create a suction. Place a piece of roll sulfur in a small evaporating dish or crucible on the wire gauze. Ignite the sulfur with the Bunsen flame. Raise the ring holding the wire gauze until it touches the funnel. Using a fishtail burner, gently heat the glass tube containing a coil of thin platinum wire (or some platinized asbestos or some ferric oxide). After a few minutes when the flask is full of white fumes, disconnect it from the apparatus and shake it until the white fumes dissolve. Again connect up the apparatus and fill the flask again with the white fumes. Dissolve these white fumes as before. Repeat 3 or 4 times.

(b) Test some of the solution in the flask with litmus paper. Result?

(c) Half fill a test tube with some of the solution from the flask. Add some barium chloride solution. Result? This is a

test for the sulfate (SO_4^{--}) ion. To confirm add some concentrated hydrochloric acid. A precipitate of barium sulfate is not affected; other possible precipitates dissolve.

Caution: Do not get sulfuric acid on your skin or clothing. Your instructor will advise you how to dispose of the sulfuric acid remaining after experiments.



2. Pour a test-tubeful of water into a beaker. **At arm's length** slowly pour $\frac{1}{4}$ of a test-tubeful of concentrated sulfuric acid into the water. **(Never pour water into sulfuric acid!)** Touch the outside of the beaker with your finger. Result?

3. Place a wood splint into a test tube containing a little concentrated sulfuric acid. After one minute remove the splint and examine it.

4. To $\frac{1}{10}$ of a test-tubeful of cane sugar standing in a test tube rack, add a few drops of concentrated sulfuric acid. **(Stand at arm's length.)** Result?

5. To $\frac{1}{6}$ of a test-tubeful of dilute sulfuric acid, add a small piece of zinc. Any action? Bring a burning wood splint to the mouth of the test tube. Result? Repeat using copper.

6. To $\frac{1}{6}$ of a test-tubeful of concentrated sulfuric acid, add a small piece of zinc. Any action? Cautiously warm the test tube. Test for hydrogen with a burning splint. Result? Cautiously note the odor of any gas evolved. Repeat using copper. Results?

Copy the following in your notebook, supplying the word or words needed to make each a complete statement. Underline the words you supply.

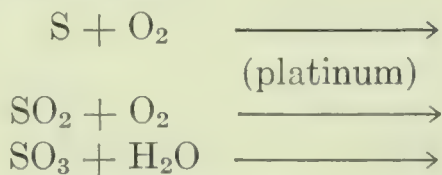
Observations and Questions on Experiments

1. The solution in the flask changes litmus from —?— to —?— indicating the presence of —?— ions. The solution contains the substance —?—.

2. The anhydride of sulfuric acid is —?—. This is the white fumes formed in Exp. 1.

3. Sulfur is burned to form —?—. This is further oxidized with the aid of —?— as a catalyst to form white solid particles of —?—. This is dissolved in water to form —?—.

4. Complete the equations:



5. When a solution of barium chloride is added to a solution of sulfuric acid, a —?— is formed which is (soluble, insoluble) —?— in hydrochloric acid.

6. Complete the equation: $\text{BaCl}_2 + \text{H}_2\text{SO}_4 \rightarrow \text{—?—}$. This reaction is irreversible and goes to completion because —?—.

7. When sulfuric acid is mixed with water in a beaker, the outside of the beaker feels —?— indicating the evolution of —?—.

8. Sulfuric acid should be poured into water because, if water were poured into sulfuric acid, the —?— generated is concentrated near the —?— and the resultant —?— would shoot up violently and spatter the sulfuric acid. On the other hand when sulfuric acid is poured into water, the sulfuric acid being —?— than the water will —?— and the —?— generated will be —?—.

9. A wood splint is —?— by concentrated sulfuric acid.

10. Sugar is —?— by concentrated sulfuric acid.

11. Both wood and sugar are carbohydrates and contain the three elements —?— in combination. The concentrated sulfuric acid acts as —?— agent and removes the elements —?— and —?— from these substances leaving the elements —?—.

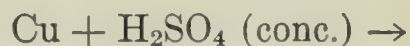
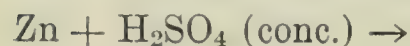
12. Dilute sulfuric acid (does, does not) —?— react with zinc

and (does, does not) —?— react with copper. Write an equation for the reaction that takes place.

13. Zinc reacts (more, less) —?— vigorously with concentrated sulfuric acid than with dilute sulfuric acid.

14. When hot concentrated sulfuric acid reacts with metals, —?— gas instead of the usual —?— gas is evolved because hot concentrated sulfuric acid acts as an —?— agent.

15. Complete the equations:



Conclusions

1. In the contact process (the most modern process) for making sulfuric acid, —?— is burned to produce —?—. This —?— is then oxidized in the presence of the catalyst —?— to produce —?—. This —?— is then dissolved in —?— to make —?—.

2. In testing for a sulfate, a solution of —?— is added to the solution to be tested. A —?— which is insoluble in —?— indicates the presence of a sulfate.

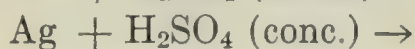
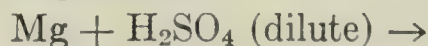
3. When dilute, sulfuric acid acts as a typical —?—, changes litmus from —?— to —?—, has a —?— taste and reacts with metals more active than hydrogen to form —?— and —?—.

4. When concentrated, sulfuric acid acts as a —?— agent and as an —?— agent.

Supplementary Exercises

1. Four physical properties of pure sulfuric acid are: color, —?—; physical state, —?—; boiling point, —?—; specific gravity, —?—.

2. Complete the equations:



3. Two experiments which can be used to distinguish dilute sulfuric acid from concentrated sulfuric acid are: —?—.

4. Complete the equations:



5. The —?— action of concentrated sulfuric acid is utilized when gasses are bubbled through it to remove —?—.

6. Three reasons for using sulfuric acid in preparing other acids are —?—.

7. Four uses for sulfuric acid other than those mentioned above are: —?—.

8. Sulfuric acid is commonly called —?—.

9. Iron receptacles can be used in shipping concentrated sulfuric acid but not for shipping dilute sulfuric acid because —?—.

10. Hot sulfuric acid causes such frightful burns on the skin because —?—.

11. Sulfuric acid has been called the "king of chemicals" because —?—.

12. If you write with a stirring rod moistened with dilute sulfuric acid on a piece of white paper and then warm the paper near a Bunsen flame, you will observe —?— because —?—.

Optional Questions

1. Commercial sulfuric acid is not colorless. What might cause the discoloration?

2. Describe the lead chamber process of making sulfuric acid.

3. Compare the action of dilute and concentrated hydrochloric acid, dilute and concentrated sulfuric acid, and dilute and concentrated nitric acid on copper. On zinc.

4. Describe the use of sulfuric acid in a particular industry.

5. In the contact process, sulfuric acid is used in the manufacture of sulfuric acid. Explain how. What is fuming sulfuric acid?

6. What is a monobasic acid? A dibasic acid? A tribasic acid? Give an example of each.

7. Mention three substances which you made or saw made in the laboratory with the use of sulfuric acid.

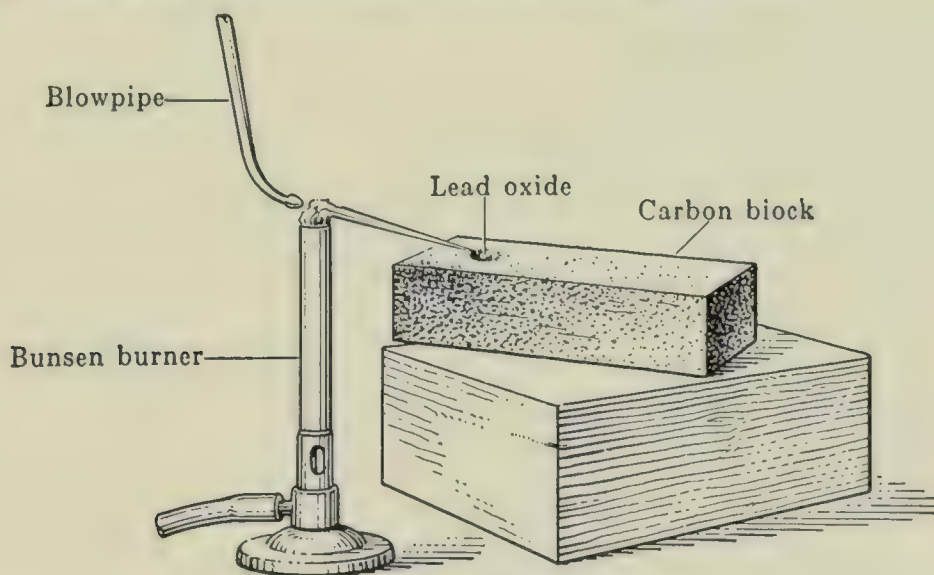
UNIT 30. CARBON

Experiments

1. *Demonstration.* Put a pinch of powdered carbon (charcoal) in a test tube. Add about $\frac{1}{2}$ a test-tubeful of water and shake. Set aside in a test tube rack. Is the carbon soluble in water? Repeat using lampblack. Repeat using solvents such as sodium hydroxide, ammonium hydroxide, hydrochloric acid, nitric acid, and an organic solvent such as carbon tetrachloride. Is carbon soluble in any of these?

2. Fill a test tube to a depth of one inch with powdered charcoal. Add a few drops of a dilute hydrogen sulfide solution. Note the odor. Cork the tube and shake vigorously. Set it aside for about 15 minutes. Now note the odor.

3. Fill a test tube to a depth of one inch with powdered boneblack. Add about 10 cc. of a solution of brown sugar. Heat the mixture. Filter the warm mixture through a piece of filter paper lined with boneblack. Note color of filtrate.



Reducing Litharge with Carbon

4. Carve out a small cavity in a charcoal block. Place a little litharge (PbO) in the cavity. Using a blowpipe, heat this PbO in contact with the charcoal (carbon) until a metallic globule appears. NOTE: Other oxides of lead may be used.

Copy the following in your notebook, supplying the word or words needed to make each a complete statement. Underline the words you supply.

Observations and Questions on Experiments

1. Carbon is (soluble, insoluble) —?— in water, —?— in alkalies, —?— in acids, and —?— in organic solvents.

2. This property of —?— helps make carbon useful in —?— and —?—.

3. The odor of the hydrogen sulfide is —?— by the powdered charcoal.

4. The molecules of hydrogen sulfide gas adhere to the —?— of the charcoal. This illustrates the process of —?— by the charcoal.

5. Filtering through boneblack causes the color of the brown sugar solution to —?—.

6. The coloring matter of the sugar —?— to the —?— of the boneblack. This illustrates the process of —?— by the boneblack.

7. When lead oxide is heated in contact with carbon, the metallic globule formed is —?—, with a —?— color and a —?— luster.

8. Complete the equation: $\text{PbO} + \text{C} \rightarrow \text{—?—}$.

9. The carbon in the reaction between lead oxide and carbon acts as a —?— agent by taking away —?— from the —?—.

Conclusions

1. An outstanding physical property of all varieties of carbon is its —?— in all of the common solvents.

2. Amorphous charcoal and boneblack are useful in removing —?— and —?— by the process of —?—. In this process the porous nature of these forms of carbon exposes a —?—.

3. Carbon is of great use as a —?— agent in extracting —?— from oxide ores and thereby obtaining valuable metals.

Supplementary Exercises

1. Charcoal tablets are sometimes prescribed for indigestion because —?—.

2. Charcoal made from fruit pits is used in gas masks because —?—.

3. It is difficult to remove stains made by printers' ink because —?—.

4. The name "lead" pencil is misleading because —?—.

5. Of the three allotropic varieties of pure carbon there are two crystalline forms, namely, —?— and —?—, and one amorphous form, namely, —?—.

6. Make a table giving the name, the method of obtaining, and two uses of 6 different commercial varieties of uncombined carbon.

7. At ordinary temperatures carbon is (active, inactive) —?— chemically. At —?— temperatures carbon combines with many metals and nonmetals.

8. Water gas is made by passing —?— over heated —?—. The equation for this reaction is: —?—.

9. Copy and complete the following equations:



10. In making silicon carbide, commonly known as —?— and very useful as an —?—, carbon is caused to react with —?—, commonly known as sand, at the high temperature of the —?— furnace. The equation for making silicon carbide is: —?—.

11. In making calcium carbide which is very useful in making the fuel gas —?—, carbon is made to react with —?— at the high temperature of —?— furnace. The equations for the preparation of CaC_2 and the reaction of CaC_2 with water are: —?—.

12. When carbon is heated in sulfur vapor, —?— is formed which is useful as —?—.

13. The atomic number of carbon is 6. The atom of carbon has a nucleus containing —?— free protons and —?— neutrons. The total number of planetary electrons is —?—. The number of electrons in the outermost ring of the carbon atom is —?—. Carbon is called an amphoteric element because it can —?— its electrons in the outermost ring. The valence of carbon is —?—.

14. A carbon atom may borrow or lend or —?— electrons with other atoms. This tendency to —?— electrons, especially with other atoms of carbon, makes possible the enormous number of carbon compounds and also explains the fact that these carbon compounds are (polar, nonpolar) —?—.

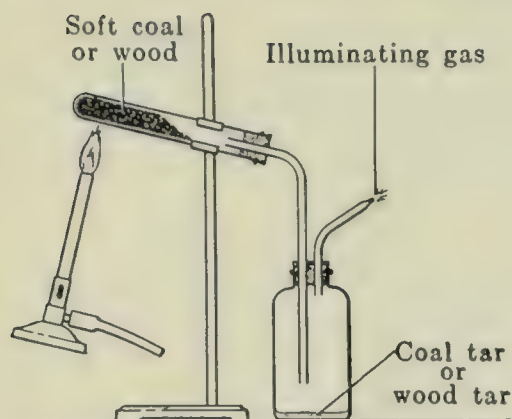
15. The study of carbon compounds is a special branch of chemistry called —?—.

Optional Questions

1. Explain the formation of coal in nature.
2. Explain the differences in the composition, properties, and uses of the various kinds of coal.
3. Explain the use of carbon in the steel industry.
4. Name two reducing agents other than carbon. Write equations to illustrate these reducing agents.
5. Describe the operation of a gas mask.
6. What is the composition of black gunpowder?
7. How would you distinguish chemically between powdered charcoal and manganese dioxide?
8. What is meant by saying that a certain sample of coal has a value of 15,000 B.T.U.?
9. Explain the Bergius process for the hydrogenation of coal.
10. What is the difference between lampblack and carbon black? Why?

UNIT 31. DESTRUCTIVE DISTILLATION

Experiments



Apparatus for the Destructive Distillation of Coal or Wood

1. Soft coal.

(a) Half fill a hard glass test tube with crushed soft coal. Arrange apparatus as shown in the figure at the left. First heat gently and then very strongly.

(b) As gas issues from the jet tube, test it with moistened litmus paper. Result?

(c) Test the gas with moistened lead acetate paper. Result?

(d) Bring a flame to the end of the jet tube. Does the gas burn?

(e) Examine the contents of the condenser bottle.

(f) When no more gas is evolved, allow the test tube to cool, break it open, and examine the residue.

2. Wood.

(a) Fill a hard glass test tube with wood splints. Use the same apparatus as in Exp. 1 above. First heat gently and then very strongly.

(b) Test the gas issuing from the jet tube with moistened litmus paper. Result?

(c) Ignite the volatile matter issuing from the mouth of the jet tube. Does it burn?

(d) Continue to heat until all the volatile matter is driven off. Allow the test tube to cool and examine the contents.

Copy the following in your notebook, supplying the word or words needed to make each a complete statement. Underline the words you supply.

Observations and Questions on Experiments

1. The coal does not burn when heated in Exp. 1 because ____?____.

2. In the process of destructive distillation, the soft coal is heated in the —?— of air to drive off the —?— matter. Part of the vapors driven off are —?— to form the liquid coal tar in the bottle.

3. The gas issuing from the soft coal changes moistened litmus from —?— to —?—. This shows the presence of a (volatile, nonvolatile) —?— (acid, base) —?—. The substance whose presence is thus shown is —?—.

4. The gas from soft coal turns lead acetate paper —?—. This shows the presence of —?— in the gas.

5. The gas from soft coal (burns, does not burn) —?—. It is commercially known as —?— gas.

6. The liquid coal tar has a —?— color and a —?— odor.

7. The solid residue obtained from the soft coal is called —?— and differs from the original coal in that —?—.

8. The wood does not burn when heated in Exp. 2 because —?—.

9. In the process of destructive distillation, the wood is heated in the —?— of air to drive off the —?— matter. Part of the vapors driven off are —?— to form the liquid wood tar.

10. The gas issuing from the jet tube in the destructive distillation of the wood changes the moistened litmus paper from —?— to —?—. This shows the presence of a (volatile, non-volatile) —?— (acid, base) —?—. The substance whose presence is thus shown is —?—.

11. The issuing gas from the destructive distillation of wood (burns, does not burn) —?—.

12. The liquid wood tar has a —?— color and a —?— odor.

13. The solid residue obtained from the destructive distillation of wood is called —?—. It differs from the original wood in that —?—.

Conclusions

1. Destructive distillation is the heating of —?— substances in the —?— of air to —?— them. In this process various —?— materials are driven off and then some of them are —?— from the —?— to the —?— state.

2. In the process of destructive distillation, soft coal leaves a residue called —?— and wood leaves a residue called —?—. The condensed liquid coal tar contains several valuable substances such as —?—. The condensed liquid wood tar is a source for several valuable substances such as —?—.

Supplementary Exercises

1. Coal was formed in nature by the partial —?— of —?— matter buried in the earth.

2. Soft coal, also called —?— coal, contains (more, less) —?— uncombined carbon and (more, less) —?— volatile carbon compounds, than hard coal (—?— coal).

3. Two uses for coke are: —?—.

4. Coal gas is purified by removing —?— and —?— before it is ready for use as a fuel gas.

5. Coal gas contains the two substances —?— and various hydrocarbon gases.

6. The poisonous ingredient of coal gas is —?—.

7. The ingredients of coal gas that burn to provide a luminous flame are —?—.

8. Two uses for charcoal are: —?—.

9. It can be shown that ordinary coke and charcoal are not pure forms of carbon by —?—.

10. Boneblack is made by the —?— of —?—. An important use of boneblack is —?—.

Optional Questions

1. How is water gas manufactured? Write the equation. How is it enriched? Why?

2. How is producer gas manufactured? What is the composition of producer gas? What kind of coal is converted into producer gas?

3. What is a calorie? What is a B.T.U.?

4. Describe the manufacture of graphite from anthracite coal.

5. What are the objections to charcoal filters for drinking water?

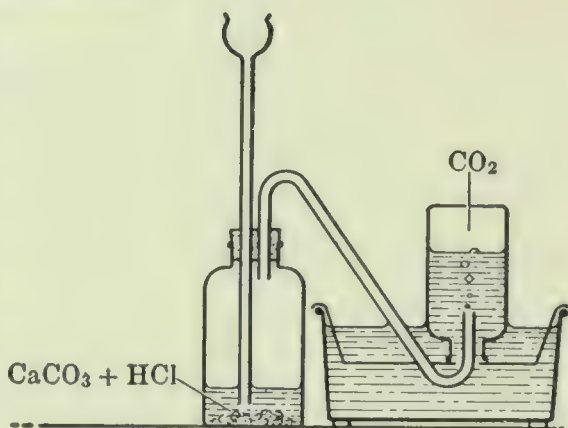
6. What is meant by adsorption? What is activated charcoal?

7. What is natural gas? How is it obtained? For what is it used?

UNIT 32. CARBON DIOXIDE

Experiments

1. Arrange apparatus as shown in the figure at the right. Place marble chips (calcium carbonate) in the generating bottle to a depth of about an inch. Add enough dilute hydrochloric acid to cover the marble. Collect 3 bottles of gas. If needed, add more acid.



Apparatus for Preparing Carbon Dioxide

NOTE: After the experiment is completed, rinse the marble chips with water and place in a container supplied by the instructor.

2. Remove the glass plate from a bottle of carbon dioxide. Any odor? Any color?

3. Stand a small candle in a wide mouth bottle. Ignite the candle. Pour some carbon dioxide from a bottle (just as you would water) over the burning candle. Result?

4. Add water to a bottle of carbon dioxide until it is $\frac{1}{3}$ full. Place the palm of your hand over the mouth of the bottle and shake vigorously. Test the liquid in the bottle with litmus. Result?

5. For at least 3 minutes, bubble some carbon dioxide into a test tube $\frac{1}{2}$ full of limewater (calcium hydroxide). Results?

6. Burn a wood splint in a bottle of air. Remove the splint. Then add limewater and shake. Result?

7. Fill a test tube $\frac{1}{2}$ full of limewater. Using a right angle glass delivery tube, bubble your breath through the limewater for about a minute. Result?

Observations and Questions on Experiments

1. The carbon dioxide is collected by the method of —?—. The gas is collected by this method, even though somewhat soluble in water, because —?—.

2. In the preparation of carbon dioxide the reaction may be written in two stages. Complete the equations:



3. Any of the common acids may be used in this method of preparing carbon dioxide because they can all supply the —?— needed and they are (less, more) —?— volatile than carbon dioxide.

4. Some properties of carbon dioxide here observed are: it has —?— odor and —?— color, it is (lighter, heavier) —?— than air, it (is, is not) —?— combustile, it (does, does not) —?— support combustion, and it is (soluble, insoluble, fairly soluble) —?— in water.

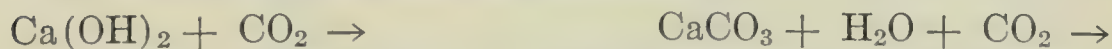
5. When the solution of carbon dioxide in water is tested with litmus, the color of the litmus changes from —?— to —?—, showing the presence of —?— ions.

6. Complete the following equation: $\text{CO}_2 + \text{H}_2\text{O} \rightarrow \text{—?—}$.

7. Because of the properties of its water solution, carbon dioxide is called an —?— anhydride.

8. When carbon dioxide is passed into limewater, there is first formed a white precipitate of —?—. With excess of carbon dioxide this white precipitate is —?—. This formation of a precipitate with limewater is used as a test for carbon dioxide.

9. Copy and complete the equations:



10. When wood burns, one of the products of combustion as shown by the limewater test is —?—. This shows that wood contains the element —?—.

11. When breath is bubbled through limewater, the result is —?—. This shows that exhaled breath contains some —?—.

Conclusions

1. Carbon dioxide is prepared in the laboratory by the reaction between —?—. In the resulting double replacement, —?— is formed which immediately decomposes into —?— and —?—.

2. Four physical properties of carbon dioxide observed in these experiments are: —?—.

3. The outstanding chemical property of carbon dioxide is its ability to react with —?— to form a weak, unstable acid, namely, —?—.

4. To test for carbon dioxide, we bubble it through —?—. A —?— results.

Supplementary Exercises

1. Carbon dioxide is obtained commercially by —?—.

2. Four important uses for carbon dioxide are: —?—.

3. Carbon dioxide is put into the air by the following three natural processes: —?—.

4. Carbon dioxide is removed from the air by green plants by the process of —?—.

5. The approximate percent by volume of carbon dioxide in air is —?—.

6. Dry ice, used as a refrigerant, is the substance —?—.

7. In the chemical fire extinguisher (ordinary type), carbon dioxide is generated by the reaction between —?—. Carbon dioxide helps put out a fire by —?—.

8. The two ingredients of a common type of baking powder that react to liberate carbon dioxide are —?—.

9. Start with the element carbon and show by means of equations that it is an acid-forming element.

10. Nonmetals are usually (acid-forming, base-forming) —?— elements, whereas metals are usually (acid-forming, base-forming) —?— elements.

11. Oxides of nonmetals (*e.g.*, CO_2 , SO_2 , N_2O_5) when dissolved in water produce —?—, and are called —?— anhydrides.

12. To test for the presence of a carbonate in a rock or stone, —?— is added to a small sample. Effervescence and the liberation of —?— gas, which turns limewater milky, indicates the presence of a carbonate.

13. In limestone regions the formation of underground caves and passages is caused by the solvent action of water and —?—

on the limestone to form soluble —?—. This substance dissolved in water makes water hard.

14. Carbon monoxide can be made by passing —?— over heated carbon. This reaction occurs in a coal furnace. The equation is —?—.

15. Two uses for carbon monoxide are: —?—.

16. Carbon monoxide is (combustible, noncombustible) —?—.

17. Carbon monoxide is a deadly poison because, if inhaled, it combines with the —?— of the blood to form a stable compound and prevents —?—.

18. Carbon monoxide is especially dangerous because it —?— odor.

19. Two sources of escaping carbon monoxide are: —?—.

20. Carbon monoxide can be oxidized to —?— at low temperatures in the presence of a catalyst called —?—.

Optional Questions

1. Explain why carbonic acid is a weak acid? an unstable acid?

2. Write the two possible equations to show the effect of passing carbon dioxide into a solution of sodium hydroxide.

3. How could you distinguish chemically between carbon monoxide and carbon dioxide? Between carbon monoxide and hydrogen?

4. Write equations to show the effect of heat on magnesium carbonate, calcium carbonate, and zinc carbonate.

5. Write an equation to show the reducing action of carbon monoxide on ferric oxide.

6. Explain with the aid of an equation how carbon monoxide can be made from formic acid. From oxalic acid.

7. Mention 3 important gaseous fuels that contain large amounts of carbon monoxide. Briefly describe how each of these fuels is manufactured.

8. Explain how the presence of carbon monoxide gas in tunnels such as the Holland Tunnel under the Hudson River can be detected.

UNIT 33. USES OF CARBON DIOXIDE

Experiments

1. Baking powders.

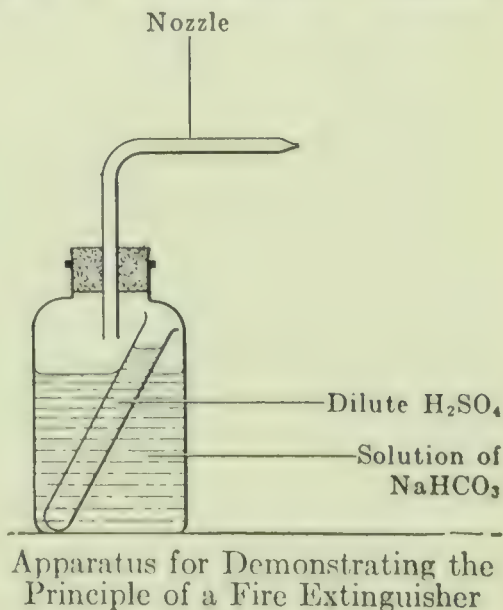
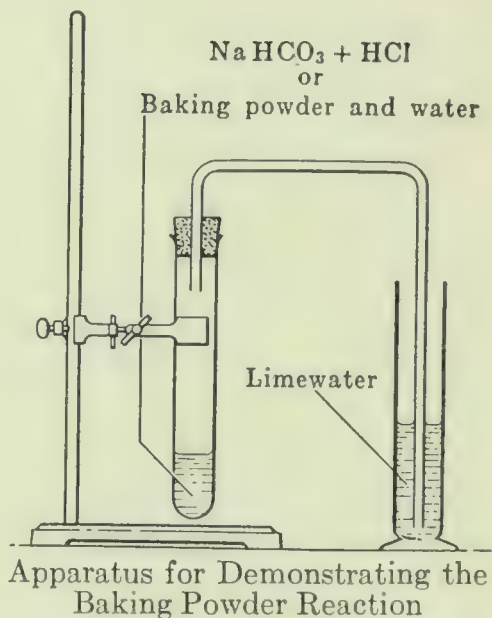
(a) Set up apparatus as shown in the figure at the right. To a little sodium bicarbonate in a test tube add a few drops of dilute hydrochloric acid. Pass some of the gas evolved through limewater. Result?

(b) Place a little cream of tartar (potassium hydrogen tartrate) on a sheet of paper. Touch with dry pieces of blue and red litmus paper. Result? Then moisten the litmus paper with water and test again. Result?

(c) On a sheet of paper mix thoroughly a little sodium bicarbonate and a little cream of tartar. This mixture contains the essential ingredients of a good baking powder. Put this mixture in a dry test tube. Any evidence of action? Use the apparatus of Exp. 1 above. Add a little water to the mixture. Result? Pass some of the gas evolved through limewater. Result?

(d) Put a little baking powder (various commercial brands should be used throughout the class) in a dry test tube. Use the apparatus of Exp. 1 above. Add a little water to the baking powder. Pass some of the gas through limewater. Result?

2. Chemical fire extinguisher. Arrange apparatus as shown in the figure at the right. Almost



fill the wide mouth bottle with a saturated solution of sodium bicarbonate. Almost fill the small test tube with dilute sulfuric acid (1:4). Stand the test tube in the bottle. Place the stopper containing the delivery tube *firmly* in the bottle. Exerting pressure with your fingers to keep the stopper from coming out, point the end of the delivery tube at the sink. Invert the bottle. Result?

Copy the following in your notebook, supplying the word or words needed to make each a complete statement. Underline the words you supply.

Observations and Questions on Experiments

1. When hydrochloric acid is added to sodium bicarbonate, the gas evolved turns limewater —?— showing that the gas is —?—.

2. Copy and complete the equation: $\text{NaHCO}_3 + \text{HCl} \rightarrow$

3. Cream of tartar will change moist litmus from —?— to —?— when —?—. This shows that cream of tartar when —?— will produce —?— ions.

4. Sodium bicarbonate and cream of tartar will react only in the presence of —?— because —?—.

5. The reaction between sodium bicarbonate and cream of tartar evolves —?— gas as shown by —?—.

6. Copy and complete the equation: $\text{NaHCO}_3 + \text{KHC}_4\text{H}_4\text{O}_6 \rightarrow \text{KNaC}_4\text{H}_4\text{O}_6 +$ —?—.

7. The commercial baking powder (name, —?—), when mixed with —?—, liberated —?— gas.

8. The liquid (mostly water) ejected from the fire extinguisher is under pressure because of the —?— volume of —?— formed by the reaction between —?— and —?—.

9. Copy and complete the equation: $\text{NaHCO}_3 + \text{H}_2\text{SO}_4 \rightarrow$

10. The function of the sodium bicarbonate in the fire extinguisher is —?—.

11. The function of the sulfuric acid in the fire extinguisher is —?—.

12. The water ejected puts out a fire by —?— and —?—.

13. The water ejected being under pressure is useful in fire fighting because —?—.

Conclusions

1. All baking powders contain the substance —?— which serves as the source of the leavening agent —?—.

2. All baking powders contain a solid substance which gives a mild —?— reaction in water. The function of this ingredient is —?—.

3. Most baking powders contain —?— or —?— which serve to keep the ingredients dry until used.

4. The chemical type of fire extinguisher generates —?— gas by the reaction between —?— and —?—. This gas serves the double purpose of —?—.

Supplementary Exercises

1. The common name for sodium bicarbonate is —?—.

2. The various types of baking powders are classified in accordance with the kind of —?— used.

3. Three types of baking powders are: —?—.

4. Baking powders are kept in closely sealed metal cans to avoid —?—.

5. In mixing cake ingredients, the baking powder should be mixed with flour (before, after) —?— adding water because —?—.

6. Sour milk can liberate —?— gas from sodium bicarbonate because the sour milk contains —?—.

7. Yeast leavens bread by causing the fermentation of —?—. This action evolves —?—.

8. The leavening action of yeast is (faster, slower) —?— than the baking powder action because —?—.

9. The chemical type of fire extinguisher is not effective against oil fires because —?—.

10. In the firefoam method of extinguishing the fire in a burning oil tank, a light foamy mixture of —?— gas and licorice extract is sprayed on the surface of the burning oil. It puts out the fire

by —?—. The —?— gas is generated in this method by the reaction between a solution of —?— (also containing the licorice extract) from one hose and a solution of —?— from another hose.

11. The small hand pump type of fire extinguisher (such as the one using Pyrene) carried on automobiles contains the substance —?—. This substance when sprayed on the fire puts it out by —?—.

12. Neither the chemical nor the firefoam types of fire extinguishers should be used on burning electrical apparatus because —?—.

13. Dry ice is the —?— form of —?—.

14. Three beverages that contain dissolved carbon dioxide are: —?—.

Optional Questions

1. Compare water, carbon dioxide, carbon tetrachloride, and foamite with respect to their ability to put out fires.

2. Give the names, formulas, and uses of three important carbonates.

3. Describe the carbon dioxide-oxygen cycle.

4. Explain the use of yeast in breadmaking.

5. Why is soda water so called?

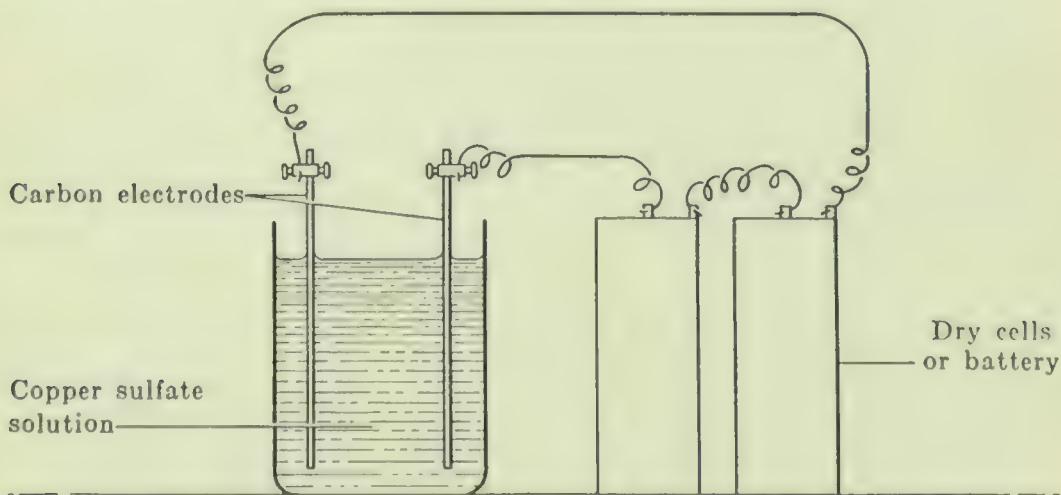
6. Describe the manufacture of dry ice. Explain its advantages over ordinary ice.

UNIT 34. METALLURGY

Experiments

1. Reduction with carbon. Carve out a small cavity in a charcoal block. Place a small quantity of litharge (PbO) in the cavity. With the aid of a blowpipe, heat with a reducing flame until metallic globules form. (The reducing flame is obtained by using a small yellow Bunsen flame. The blowpipe is placed just outside the flame, and the breath is blown gently and steadily to form a yellow cone of flame.)

2. Roasting in air followed by reduction with carbon. Place some lead sulfide (PbS) (if not available, use copper sulfide, CuS) in a porcelain crucible. Heat strongly for a few minutes. Cautiously smell the gas evolved. The crucible now contains lead oxide. (If CuS is used, the crucible contains CuO and reddish-brown Cu_2O .) Repeat Exp. 1 above, using some of the contents of the crucible instead of the litharge. Result?



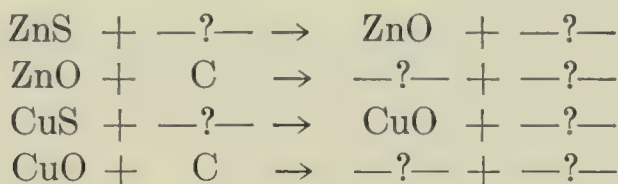
Apparatus for the Electrolysis of Copper Sulfate

3. Electrolysis. Arrange apparatus as shown in the figure above. Place a solution of copper sulfate in the glass tumbler. The electrodes are carbon rods. Use 2 dry cells in series as the source of current and send electricity through the solution for a few minutes. Examine both electrodes. Result?

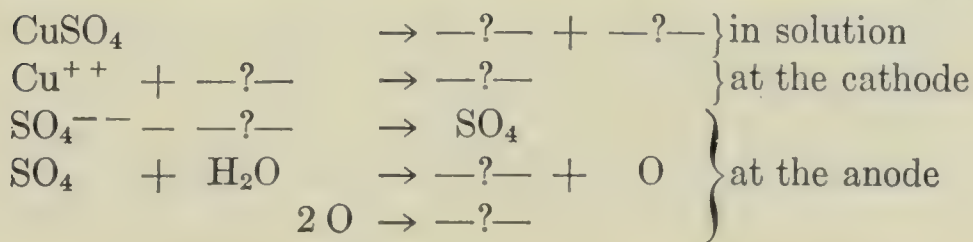
Copy the following in your notebook, supplying the word or words needed to make each a complete statement. Underline the words you supply.

Observations and Questions on Experiments

1. Copy and complete the equation: $\text{PbO} + \text{C} \rightarrow$
2. When lead oxide is heated with carbon, the carbon acts as —?— agent by taking the carbon away from the lead oxide and combining with this —?— to form —?—.
3. When an ore is roasted in air, it reacts with the —?— of the air.
4. When zinc sulfide (copper sulfide) ore is roasted in air, —?— gas with its —?— odor is evolved.
5. Copy and complete the equations:



6. When an electric current is sent through a solution of copper sulfate, —?— is deposited on the —?— electrode.
7. When copper sulfate is dissolved in water, —?— ions and —?— ions are formed. When the electricity is sent through, the —?— ions with their $++$ charge are attracted to the —?— electrode and there (gain, lose) —?— electrons to form —?— of copper. Correspondingly the —?— ions with their $--$ charge are attracted to the —?— electrode and there (gain, lose) —?— electrons to form the —?— group which immediately reacts with water to form —?— and —?—.
8. Complete the following ionic and electronic equations:



Conclusions

1. Metals are usually extracted from metallic oxide ores by removing the —?— by means of —?— which serves as —?— agent. The process is called —?—.

2. If an ore occurs as a metallic sulfide, the ore is first —?—, thus converting the sulfide into the —?—. This —?— is then reduced with —?—.

3. In the electrolytic method, electricity is sent through the ionized ore (either in solution or in the melted state) and the metal is deposited and collected at the —?—.

Supplementary Exercises

1. Complete the following equations which show how zinc can be extracted from zinc carbonate ore:



2. Carbon is used as the reducing agent in removing oxygen from many metallic oxide ores because —?—.

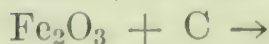
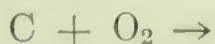
3. Five common metals extracted from their oxide ores by reduction with carbon are: —?—.

4. Four metals extracted from their ores by electrolytic methods are: —?—. These metals cannot be extracted from their ores by reduction with carbon because —?—.

5. The following three metals are found in the free or uncombined state in nature and are extracted by simple mechanical means: —?—.

6. (a) Draw a labeled diagram of a blast furnace to show the extraction of iron from hematite, Fe_2O_3 . (A commercial reduction method.)

(b) Copy and complete these equations for reactions that occur in the blast furnace.



(c) A flux is added to the charge in a blast furnace in order to —?—.

(d) The usual flux in a blast furnace is the substance —?— because the usual impurity is —?—.

(e) The slag in a furnace is the result of —?—.

(f) The iron coming out of the blast furnace is called —?— iron.

(g) Four impurities found in this —?— iron are: —?—.

(h) In making wrought iron, these impurities must be —?—.

(i) In making steel, these impurities must be —?— and one or more of various substances such as —?— are added to give more desirable properties.

7. (a) Draw a labeled diagram of the furnace for the extraction of aluminum from bauxite, Al_2O_3 . (A commercial electrolysis method.)

(b) Complete the equation: $\text{Al}_2\text{O}_3 \xrightarrow{\text{electricity}}$

(c) The electrolytic method of extracting aluminum from bauxite was invented by a young American named —?— in 1886.

(d) The cryolite serves the purpose of —?—.

(e) The aluminum collects at the —?— and the oxygen at the —?—.

8. Copy and complete the following table of some common ores:

NAME OF ORE	METAL OBTAINED	CHEMICAL NAME OF IMPORTANT COMPOUND IN ORE	FORMULA OF COMPOUND
Bauxite			
Cassiterite			
Chalcopyrite			
Cinnabar			
Cuprite			
Galena			
Hematite			
Limonite			
Magnetite			
Siderite			
Smithsonite			
Zinc blende			
Zincite			

Optional Questions

1. Describe the manufacture of wrought iron. Draw a diagram of the furnace used.
2. Compare wrought iron with cast iron with respect to properties and uses.
3. With the aid of a diagram describe the manufacture of steel by the Bessemer process. Explain the use of the basic lining.
4. With the aid of a diagram describe the open hearth process of making steel.
5. Describe the crucible process of making steel.
6. Describe the process of extracting copper from its sulfide ore.
7. How is mercury obtained from cinnabar?
8. Describe Parke's process for separating silver from lead.
9. Describe two ways of mining gold.
10. What steps are necessary to obtain copper from a copper sulfide ore? Write equations. Why must copper for industrial use be 100% pure?
11. With the aid of a diagram and the electron theory explain the electrorefining of copper.

UNIT 35. RELATIVE ACTIVITY OF METALS

Experiments

1. Set up a test tube rack with 6 test tubes. (Use small test tubes if available.) Half fill each of the test tubes with one of the first six solutions listed in the table below. Either number the test tubes or keep them in exact order. Place a clean strip of zinc in each test tube and allow to stand for five minutes. Test any escaping gas with a burning splint. Make a table in your notebook like the one below and record your results.

2. Repeat Exp. 1 above using strips of copper instead of zinc. Record your results in the table in your notebook.

No.	METAL	SOLUTION	NATURE OF DEPOSIT (IF ANY)	EQUATION FOR REACTION (IF ANY)
1	Zn	HCl(dil)		
2	Zn	Pb(NO ₃) ₂		
3	Zn	AgNO ₃		
4	Zn	HgNO ₃		
5	Zn	CuSO ₄		
6	Zn	Al ₂ (SO ₄) ₃		
7	Cu	HCl(dil)		
8	Cu	Pb(NO ₃) ₂		
9	Cu	AgNO ₃		
10	Cu	HgNO ₃		
11	Cu	CuSO ₄		
12	Cu	Al ₂ (SO ₄) ₃		

Copy the following in your notebook, supplying the word or words needed to make each a complete statement. Underline the words you supply.

Conclusions

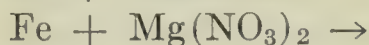
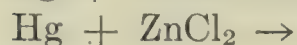
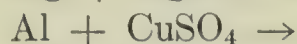
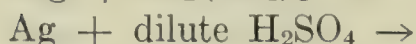
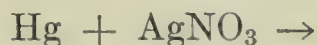
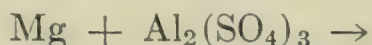
1. Zinc can replace the five metals —?— but cannot replace —?— from compounds in solution.
2. Copper can replace the two metals —?— but cannot replace the four metals —?— from compounds in solution.
3. Metals can replace other metals (and hydrogen) from compounds in solution provided the metal in the uncombined state is —?—.

Supplementary Exercises

1. Beginning with potassium and including hydrogen list 16 of the common metals in the order of their activity (electrochemical series.)
2. The metals at the (top, bottom) —?— of the electrochemical series corrode easily and are (easily, with difficulty) —?— extracted from their ores.
3. The metals at the (top, bottom) —?— of the electrochemical series do not corrode readily and are (easily, with difficulty) —?— extracted from their ores.
4. When zinc reacts with copper sulfate solution, the color of the solution changes from —?— to —?— because —?—.
5. Aluminum and zinc are (more, less) —?— active than iron and yet corrode (more, less) —?— readily in air than iron because —?—.
6. Three ingredients of air that cause corrosion of metals are: —?—.
7. Gold, silver, and copper were known and used long before iron and aluminum because —?—.
8. Aluminum is used to obtain the metal chromium from chromium oxide. Since the aluminum is (more, less) —?— active than the chromium, it will combine with the —?— and liberate

the —?—. The aluminum here acts as a —?— agent. The equation is —?—.

9. Complete the following equations. When no reaction occurs, indicate this fact.

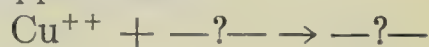


10. According to the electron theory, the metals in the electrochemical series are arranged in the order of their readiness to —?— electrons.

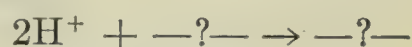
11. According to the electron theory, when a metal replaces another metal from a compound in solution, the atoms of the replacing metal (gain, lose) —?— electrons and become —?—. The ions of the other metal (gain, lose) —?— electrons and are converted from ions to —?—.

12. Copy and complete the following equations that indicate the ionic and electronic changes that take place when:

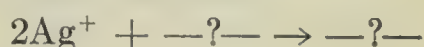
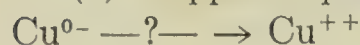
(a) Zinc is placed in a solution of copper sulfate.



(b) Zinc is placed in dilute sulfuric acid.



(c) Copper is placed in a solution of silver nitrate.



Optional Questions

1. Aluminum is more active than iron and yet iron corrodes faster. Explain.

2. What metals are insoluble in nitric acid? How can you dissolve these metals?

3. Explain why the replacement series of metals is also called the electromotive series.

4. With the aid of a diagram describe the electroplating of a brass spoon with silver.

5. Why are metallic objects chromium plated?

6. Write electronic equations to show the action of potassium with water.

UNIT 36. TESTS FOR METALS, ACID RADICALS, AND SOME COMMON GASES

Experiments

1. Flame tests. The instructor will have prepared various salt solutions in vials or bottles. (Use solutions of compounds listed in the table below.) Inserted in the cork stopper of each bottle is a clean nichrome wire which dips into the solution. Take hold of each cork stopper and hold each wire in the non-luminous Bunsen flame. Note the color of each flame. Try to find the flash of color other than yellow. Look through the blue cobalt glass for those indicated in the table. Make a table like the one below and record your results.

COMPOUND OF THE METAL	FORMULA OF COMPOUND USED	COLOR OF FLAME	COLOR SEEN THROUGH BLUE COBALT GLASS
Sodium	<i>NaCl</i>	<i>yellow</i>	
Potassium	<i>KCl</i>	<i>purple</i>	<i>purple</i>
Calcium	<i>CaCl</i>	<i>red</i>	×
Lithium	<i>LiCl</i>	<i>red</i>	×
Barium		<i>green</i>	×
Sodium and potassium	NaCl+KCl	<i>purple</i>	

NOTE: *It is suggested to the instructor that, in case of time shortage, the following experiments might be apportioned among the individual students and that the results in all tests be observed and noted by everybody. Each student should perform at least one cobalt nitrate test and at least one borax bead test.*

2. Cobalt nitrate tests. Carve out a small cavity in a charcoal block. Place a small quantity of a zinc compound in the cavity and with the aid of a blowpipe heat until red hot. Use an oxidizing flame. (An oxidizing flame is obtained by using a small yellow Bunsen flame, inserting the tip of the blowpipe in

the flame and blowing steadily to produce a bluish-green flame.) Moisten the residue with 2 or 3 drops of cobalt nitrate solution. Again heat to red heat. Allow residue to cool and note color. Repeat, using an aluminum compound in a fresh cavity. Repeat, using a magnesium compound in a fresh cavity. Make a table like the one below and record your results.

3. Borax bead tests. Place some powdered borax on a sheet of paper. Take up some borax on a loop of a platinum wire and repeatedly hold in the Bunsen flame until the borax has melted and formed a clear glassy bead. (If platinum wire is not available, the bead may be made at the end of a thin glass stirring rod.) Touch the bead, while hot, against a *tiny* particle of a cobalt compound, so that it adheres. Heat again until the bead is red hot. Observe color when cold. Repeat (using a clean wire or a clean stirring rod in each case) with compounds of chromium, nickel, and iron. (If the color is too dark, repeat, using less of the compound tested.) Make a table like the one below and record your results.

COBALT NITRATE TESTS

COMPOUND OF METAL	FORMULA OF COMPOUND USED	COLOR
Zinc		
Aluminum		
Magnesium		

BORAX BEAD TESTS

COMPOUND OF METAL	FORMULA OF COMPOUND USED	COLOR
Cobalt		
Chromium		
Iron		
Nickel		

Copy the following in your notebook, supplying the word or words needed to make each a complete statement. Underline the words you supply.

Supplementary Review of Previous Tests for Metals, Acid Radicals, and Some Common Gases

1. Hydrogen sulfide test for metals. (See Unit 27.) Hydrogen sulfide gas passed into solutions of certain metallic compounds produces a colored precipitate. Make a table giving the name of the precipitate and its color when hydrogen sulfide is

passed into compounds of (a) cadmium, (b) copper, (c) zinc, (d) antimony, (e) lead.

2. Test for an ammonium compound. (See Unit 24.) To some of the solid unknown add —?— and heat. The evolution of —?— gas, with its characteristic odor, and its ability to turn wet litmus from —?— to —?— shows the presence of an ammonium salt. The equation is $\text{NH}_4\text{Cl} + \text{—?—} \rightarrow$

3. Test for a chloride. (See Unit 15.) To the unknown in solution add some —?— solution. A —?— which turns purple in the sunlight and is insoluble in nitric acid shows the presence of a chloride. The equation is: $\text{NaCl} + \text{—?—} \rightarrow$

4. Test for a bromide. (See Unit 13.) To the unknown in solution add some —?— and —?—. Shake. An —?— color in the bottom carbon bisulfide layer shows that free —?— has been liberated from a bromide in the unknown. The equation is: $\text{NaBr} + \text{—?—} \rightarrow$

5. Test for an iodide. (See Unit 14.) To the unknown in solution add some —?— and —?—. Shake. A —?— color in the bottom carbon bisulfide layer shows that free —?— has been liberated from an iodide in the unknown. The equation is $\text{NaI} + \text{—?—} \rightarrow$

6. Test for a sulfide. (See Unit 27.) To some of the solid unknown add a few drops of —?—. The offensive odor of rotten eggs indicates the evolution of —?— gas and shows the presence of a sulfide. This gas also turns lead acetate paper —?—. The equation is: $\text{ZnS} + \text{—?—} \rightarrow$

7. Test for a sulfate. (See Unit 29.) To the unknown in solution add some —?— solution. A —?— insoluble in hydrochloric acid shows the presence of a sulfate. The equation is: $\text{Na}_2\text{SO}_4 + \text{—?—} \rightarrow$

8. Test for a sulfite. (See Unit 28.) To the solid unknown add a little —?—. The evolution of —?— gas with its suffocating odor shows the presence of a sulfite. This gas also decolorizes a dilute potassium permanganate solution. Complete the following equation: $\text{Na}_2\text{SO}_3 + \text{—?—} \rightarrow$

9. Test for a carbonate. (See Unit 32.) To some of the solid

unknown add a little —?—. The evolution of —?— gas which when passed into limewater causes it to —?— shows the presence of a carbonate. Complete the equation: $\text{Na}_2\text{CO}_3 + \text{—?—} \rightarrow$

10. Test for a nitrate. (See Unit 25.) To some of the unknown in solution add some freshly prepared —?— solution. Then carefully pour down the side of the test tube some concentrated sulfuric acid. A —?— in between the two layers indicates the presence of a nitrate.

11. Copy and complete the following table giving the tests for some common gases.

GAS	FORMULA	PROCEDURE	RESULTS
Ammonia			
Carbon dioxide			
Carbon monoxide			
Chlorine			
Hydrogen			
Hydrogen chloride			
Hydrogen sulfide			
Nitrogen			
Nitric oxide			
Nitrogen peroxide			
Nitrous oxide			
Oxygen			
Sulfur dioxide			

Optional Questions

Describe briefly the tests that you would perform and the results that should be obtained in identifying both the metallic and nonmetallic parts of each of the following:

- Potassium carbonate
- Sulfuric acid
- Zinc nitrate
- Sodium iodide
- Copper sulfate
- Calcium chloride
- Ammonium bromide
- Barium sulfide
- Aluminum sulfite
- Chromium chloride

UNIT 37. IDENTIFICATION OF AN UNKNOWN SALT

PROCEDURE: Your unknown substance will be a simple salt, that is, a compound containing a metal and an acid radical. Follow the procedures given in Unit 36 and systematically test for the presence of each metal and acid radical listed below. In view of the nature of this unknown substance you must expect all but two of the tests to produce negative results. Dissolve only one-fourth of your unknown substance in a test tube of water. Use only small portions of solution or solid for each test. If you have time, analyze more than one unknown.

Metals

Potassium, K
Sodium, Na
Calcium, Ca
Lithium, Li
Barium, Ba
Cobalt, Co
Chromium, Cr
Zinc, Zn
Aluminum, Al
Ammonium (radical), NH_4

Acid Radicals

Chloride, Cl
Bromide, Br
Iodide, I
Sulfide, S
Sulfate, SO_4
Sulfite, SO_3
Carbonate, CO_3
Nitrate, NO_3

Conclusions

Copy and complete the table below:

NO. OF UNKNOWN	METAL	ACID RADICAL	NAME OF SALT	FORMULA OF SALT

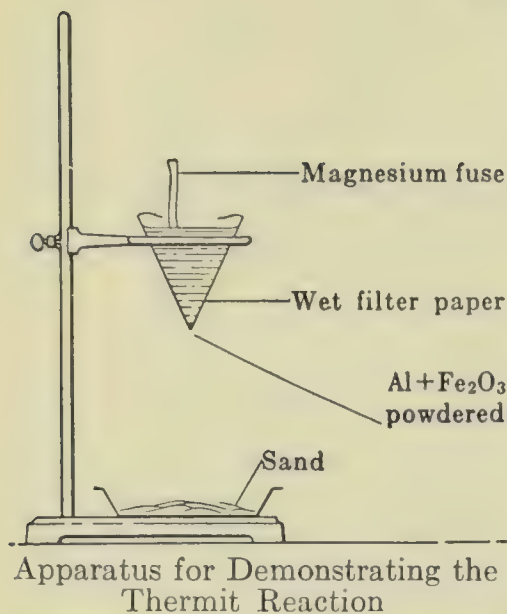
Copy and complete the following in your notebook:

Supplementary Exercises

1. We identify substances by observing —?—.
2. By qualitative analysis of substances we find —?—.
3. By quantitative analysis of substances we find —?—.
4. In spectroscopic analysis substances are —?—.
5. In X-ray analysis substances are —?—.

UNIT 38. ALUMINUM AND SOME OF ITS COMPOUNDS

Experiments



1. *Demonstration.* Thermite reaction. Thoroughly mix on a sheet of paper about 2 teaspoonfuls of ferric oxide (11 g.) with about one-half this bulk of powdered aluminum ($8\frac{1}{2}$ g.) (thermite mixture). Fold together two pieces of filter paper (as for filtering), wet with water, and arrange as in the figure at the left. Place a pan of sand underneath the ring. Pour the mixture

into the filter paper. Stick a 3-inch piece of magnesium ribbon into the thermite mixture to serve as a fuse. Cautiously ignite the magnesium. **Stand a few feet away.** Result?

2. Aluminum hydroxide.

(a) To $\frac{1}{4}$ of a test-tubeful of aluminum sulfate solution add an equal volume of ammonium hydroxide. Result? The precipitate formed is aluminum hydroxide. Note its appearance.

(b) As a coagulant. To some water in two glass cylinders or wide mouth bottles add a little fine clay and stir until the water is turbid. Set one cylinder aside. To the other cylinder add $\frac{1}{2}$ a test-tubeful of aluminum sulfate solution and stir. Then add slowly without stirring a test-tubeful of ammonium hydroxide. Allow both to stand for about 15 minutes. Result? Compare the two cylinders.

(c) As a mordant. Place a small piece of white cotton cloth into an evaporating dish one-third full of logwood solution (or similar dye). Boil for a minute. Remove the cotton and wash it thoroughly with water and dry. Result?

Now mordant a piece of cotton cloth by first soaking in

aluminum sulfate solution (squeeze out the excess liquid) and then placing in an ammonium hydroxide solution. Again squeeze out the excess liquid. Place the mordanted cloth in logwood solution and boil for a minute. Wash the cloth with water and then dry it. Compare the mordanted piece of cloth with the unmordanted piece.

Copy the following in your notebook, supplying the word or words needed to make each a complete statement. Underline the words you supply.

Observations and Questions on Experiments

1. The reaction between ferric oxide and aluminum is decidedly (endothermic, exothermic) —?—.
2. In the thermit reaction, —?— in the —?— state is produced. This drops through the filter paper into the sand below.
3. The magnesium placed in the thermit mixture serves the purpose of —?—.
4. The aluminum in the thermit reaction acts as —?— agent. Aluminum is a (more, less) —?— active metal than iron.
5. The reaction between ferric oxide and aluminum is called the thermit reaction because the —?—.
6. The equation for the thermit reaction is: $\text{Fe}_2\text{O}_3 + \text{Al} \rightarrow$ —?—.
7. Aluminum hydroxide may be described as a —?— precipitate.
8. Complete the equations:

$$\text{Al}_2(\text{SO}_4)_3 + \text{Ca}(\text{OH})_2 \rightarrow \qquad \qquad \text{Al}_2(\text{SO}_4)_3 + \text{NH}_4\text{OH} \rightarrow$$
9. The precipitation of —?— in water made turbid with clay particles causes —?—.
10. Aluminum hydroxide can be deposited as a precipitate in the fibers of a piece of cloth by first dipping the cloth into some —?— solution and then into some —?— solution.
11. Cloth impregnated with a precipitate such as aluminum hydroxide is called —?— cloth. Cloth is thus —?— in order to make dyes such as logwood —?—.

12. The combination of a precipitate such as aluminum hydroxide and a dye such as logwood is known as a —?—.

Conclusions

1. In the thermit process, aluminum acts as a —?— agent. The combination of aluminum with the —?— of ferric oxide evolves sufficient heat to cause the —?—.

2. Aluminum hydroxide is used to —?— and to —?—. Two properties of aluminum hydroxide that make possible these uses are: —?—.

Supplementary Exercises

1. Complete the equation: $\text{Al} + \text{HCl} \rightarrow \text{—?—}$.

In this reaction aluminum acts as a (metal, nonmetal) —?—.

2. Complete the equation: $\text{Al} + \text{NaOH} \rightarrow \text{—?—}$.

In this reaction aluminum acts as a (metal, nonmetal) —?—.

3. An element which behaves as a metal in some reactions and as a nonmetal in other reactions is called an —?— element. An element other than aluminum which behaves this way is —?—.

4. Aluminum is a (more, less) —?— active metal than magnesium. Aluminum utensils are not cleaned with alkalies because —?—.

5. Four uses of the metal aluminum other than in the thermit process are: —?—.

6. Aluminum does not corrode readily in air because it forms a —?— of —?— on its surface.

7. Aluminum is surpassed only by the metals —?— and —?— in importance and variety of uses.

8. Some important alloys of aluminum are: magnalium used for —?— and which contains —?—; aluminum bronze used for —?— and which contains —?—; and duralumin used for —?— and which contains —?—.

9. Aluminum cannot be extracted from bauxite (Al_2O_3) by reduction with carbon because —?—. Aluminum is extracted from this ore by —?— process.

10. The chemical term “alum” refers to a double —?— which

is a combination of the bivalent nonmetallic —?— group, a —?— metal, and a —?— metal. Crystals of the alums all contain —?— molecules of water of crystallization.

11. Alundum, an artificial abrasive, is similar in composition to the natural abrasive —?—, and is the chemical substance —?—.

12. Synthetic rubies and sapphires, containing traces of impurities to provide —?—, are identical in composition with the natural gems. They are the chemical substance —?—.

13. Pure clay (kaolin) is the chemical substance —?—. Common clay owes its color to —?—.

14. Cement is made by heating together —?— and —?—. Cement hardens by —?—. Concrete is made by mixing —?—. Reinforced concrete contains in addition —?—.

15. Bricks and ordinary pottery are made by adding water to impure —?— to make it —?—. This is shaped and heated in a kiln. The more expensive nonporous ceramic called porcelain is made by heating a previously shaped plastic mixture of the three substances —?—. Then a glaze is baked on it.

Optional Questions

1. Explain the use of aluminum in the steel industry.
2. Explain the use of aluminum in extracting chromium and manganese from their ores.
3. Describe the construction and operation of the photoflash bulb.
4. How are artificial rubies and sapphires made?
5. What is Fuller's earth? mica? feldspar?
6. In mordanting, cloth is very often treated with aluminum acetate and then with steam. Explain.
7. Discuss the effect of acid dyes and basic dyes on textile fabrics. When must a mordant be used?
8. Compare the use of alum and of copper sulfate in the purification of water.

UNIT 39. HARD WATER

Experiments

1. To a test tube $\frac{1}{2}$ full of distilled water add a few drops of soap solution. Shake vigorously. Are suds produced? Repeat, using tap water. Are as many suds produced?

2. Pass carbon dioxide gas into a test tube $\frac{2}{3}$ full of limewater. Continue to pass the gas into the limewater until the white precipitate first formed dissolves. The test tube now contains a solution of calcium bicarbonate, $\text{Ca}(\text{HCO}_3)_2$, in water. If the solution is not clear, filter it. Water containing calcium bicarbonate or similar salts in solution is called temporary hard water. Divide this temporary hard water into three portions.

3. To the first portion of temporary hard water add a few drops of soap solution. Shake vigorously. Any suds formed? What does form? Continue to add soap solution and shake. Do suds finally form?

4. Gently boil the second portion of temporary hard water in a test tube for a few minutes. Pass any gas evolved into some limewater in a test tube. Results? Test the temporary hard water which was boiled with some soap solution. Result?

5. To the third portion of temporary hard water add limewater drop by drop until no more precipitate forms. Filter and test the filtrate with a few drops of soap solution. Result?

6. To a test tube $\frac{1}{2}$ full of calcium sulfate solution add a few drops of soap solution. Shake vigorously. Any suds? Water containing dissolved calcium sulfate or similar compounds is called permanent hard water.

7. Boil some calcium sulfate solution (permanent hard water) for 2 or 3 minutes. Any change? Test with soap solution. Result?

8. To a test tube $\frac{1}{2}$ full of permanent hard water add $\frac{1}{4}$ of a test-tubeful of sodium carbonate solution drop by drop. Result? Filter and test the filtrate with a few drops of soap solution. Result?

Copy the following in your notebook, supplying the word or words needed to make each a complete statement. Underline the words you supply.

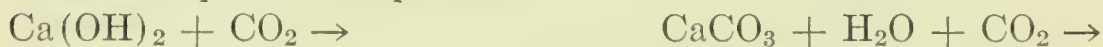
Observations and Questions on Experiments

1. When soap is shaken with distilled water, —?— are produced.

2. Our tap water produces (less, an equal amount of, more) —?— suds with soap than does distilled water.

3. The white precipitate first formed when carbon dioxide is passed into limewater is —?—.

4. Complete the equations:



5. Temporary hard water (does, does not) —?— produce suds with soap.

6. Temporary hard water reacts with soap to form a white flocculent —?—.

7. Upon continued treatment of hard water with soap solution there is produced —?—.

8. By boiling temporary hard water, there is produced —?— gas and also a precipitate of —?—.

9. The filtrate from boiled temporary hard water (does, does not) —?— produce suds with soap.

10. Complete the equation: $\text{Ca(HCO}_3)_2 \rightarrow$ —?—.

11. When temporary hard water is treated with limewater and then filtered, the filtrate (does, does not) —?— produce suds with soap.

12. Complete the equation: $\text{Ca(HCO}_3)_2 + \text{Ca(OH)}_2 \rightarrow$ —?—.

13. Water containing dissolved calcium sulfate (permanent hard water) produces a —?— instead of —?— with soap.

14. When permanent hard water is boiled, the water (is, is not) —?— softened (i. e., able to produce suds with soap).

15. When sodium carbonate (washing soda) is added to permanent hard water there is formed a —?—. When this mixture is filtered, the filtrate (is, is not) —?— soft water; it (does, does not) —?— produce suds with soap.

16. Complete the equation: $\text{CaSO}_4 + \text{Na}_2\text{CO}_3 \rightarrow \text{---?---}$.

Conclusions

1. Hard water contains (dissolved, undissolved) ---?--- salts, usually compounds of the metal ---?---. Hard water forms a ---?--- instead of ---?--- with soap.

2. Temporary hard water usually contains the compound ---?--- in solution. It is formed in nature by the action of water and ---?--- on ---?---.

3. To soften hard water (enable it to produce suds with soap) the (dissolved, undissolved) ---?--- minerals are converted to ---?--- and then filtered out.

4. Temporary hard water can be softened by ---?--- or by the addition of ---?---.

5. Permanent hard water is so called because it cannot be softened by ---?---. Permanent hard water is misnamed because it can be softened by the addition of ---?---.

Supplementary Exercises

1. Besides calcium bicarbonate, another substance which may make water temporarily hard is ---?---.

2. Besides calcium sulfate, another substance which may make water permanent hard water is ---?---.

3. Boiler scale and the deposit in tea kettles are deposits of ---?--- caused by ---?---.

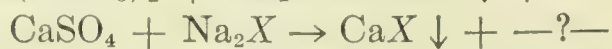
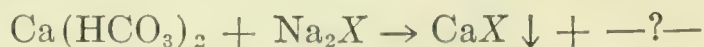
4. Underground caves and passages are formed in limestone regions by the solvent effect of ---?--- and ---?--- on the limestone.

5. Stalactites and stalagmites are formed in limestone caves by the precipitation of ---?--- from the ---?---.

6. Two substances, other than soap, that may be found in hard water soap are ---?---.

7. To soften water on a large scale a commercial product called ---?--- is often used. This is a complex sodium silicate. When hard water is passed through beds of this material, the water is softened by ---?--- the dissolved minerals.

8. Representing the formula of permutit as Na_2X , complete the equations:



9. The CaX can be changed back to the Na_2X by treatment with a ---?--- solution and so a permutit system lasts almost indefinitely.

10. Write the equation for the change of CaX back to Na_2X .

Optional Questions

1. Explain the formation of a deposit inside a teakettle. How could you remove it?

2. Explain how limestone deposits are formed in the earth.

3. How could you show that a sample of water contained dissolved minerals?

4. Write a report on the industrial methods used to soften hard water.

UNIT 40. CALCIUM COMPOUNDS

Experiments

1. Lime. Place a small piece of marble (CaCO_3) on a wire gauze and heat strongly for about 15 minutes. (Proceed to Exp. 2 in the meantime.) Allow to cool. Then place it in a test tube and add water. Any change in temperature? Test with litmus. Result?

2. Plaster of Paris. Mix about a tablespoonful of plaster of Paris, $(\text{CaSO}_4)_2 \cdot \text{H}_2\text{O}$, with enough water in an evaporating dish to make a thick paste. Place on a piece of paper. Wet a new coin with kerosene and press the coin into the paste. Set aside to harden for about 15 minutes. Remove the coin and examine the impression.

3. Mortar. Mix about a tablespoonful of quicklime, CaO , with twice its volume of fine sand in an evaporating dish. Add enough water to make a thick paste. Fill a small match box with the mixture and set aside for a few days. Then examine for appearance and hardness.

4. Cement. Mix about a tablespoonful of Portland cement with three times its volume of fine sand in an evaporating dish. Add enough water to make a thick paste. Fill a small match box with the mixture and set aside for a few days. Then examine appearance and hardness.

Copy the following in your notebook, supplying the word or words needed to make each a complete statement. Underline the words you supply.

Observations and Questions on Experiments

1. Write the equation for the heating of marble.

2. When water is added to quicklime, there is formed —?— which turns litmus from —?— to —?—. This action is commercially called —?— of lime and is accompanied by (no change, a drop, a rise) —?— in temperature.

3. Complete the equation: $\text{CaO} + \text{H}_2\text{O} \rightarrow \text{—?—}$.

4. When plaster of Paris sets or hardens, it takes on —?—

to form hard crystals of —?—. Complete the equation:



5. Mortar is made by mixing —?—. The —?— and —?— unite to form —?—. The —?— in the mixture gives the mortar extra body and strength.

6. After exposure to air for a week the mortar mixture becomes —?—.

7. In the setting or hardening of mortar the excess —?— first evaporates leaving a hard mass of —?— and —?—. Then the —?— of the air combines with this to form even harder, more insoluble —?— by the following reaction: —?—.

8. After a week the mixture in Exp. 4 becomes —?—.

9. Cement can “set” or harden under water because —?—.

Conclusions

1. Quicklime (also called lime) is made by heating —?— to drive off —?— gas.

2. When water is added to quicklime, —?— is formed. This process is called —?—.

3. Plaster of Paris sets or hardens by —?— to form hard crystals of —?—.

4. Mortar sets and hardens by the evaporation of the excess —?— and the subsequent action of the —?— of air on the —?— to form hard, insoluble —?—.

5. Cement sets or hardens by —?—.

Supplementary Exercises

1. Plaster of Paris sets (more, less) —?— quickly than mortar because —?—.

2. An important reason for using plaster of Paris in making molds or casts is that upon hardening it —?—.

3. Plaster of Paris is made by heating —?— to drive off —?—.

4. Lime is shipped in airtight containers because it —?—.

5. The two equations for the air-slaking of lime are: —?—.

6. The most abundant compound of calcium found in nature is —?—. Three natural varieties of this compound are —?—.

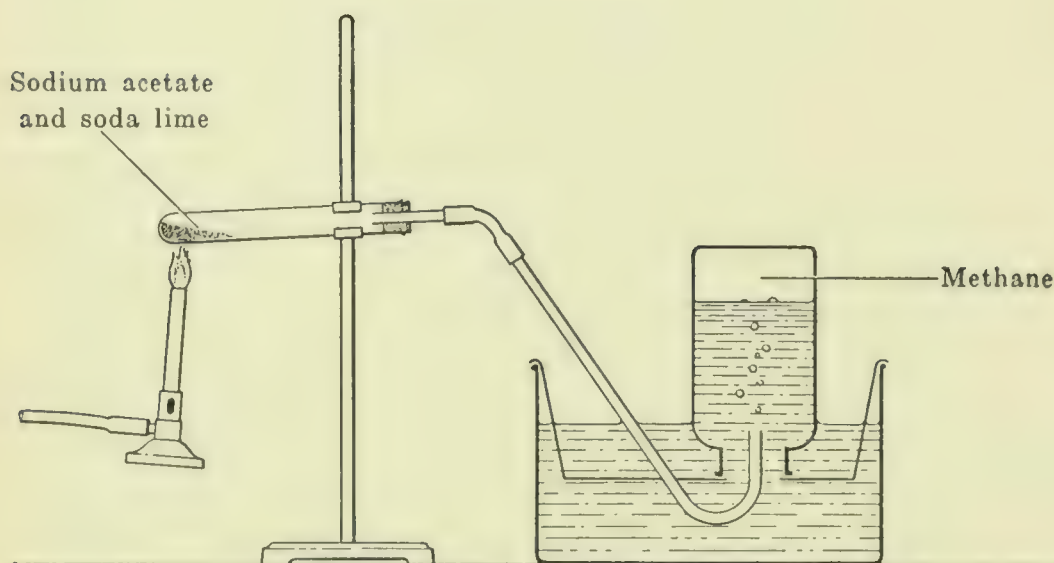
7. The metal calcium is prepared by the —?— of fused —?—.
8. Concrete is made by mixing —?—, pouring this mixture into a wood or steel mold, and then waiting until —?—.
9. Reinforced concrete contains —?—.
10. Make a table in your notebook giving the *chemical name*, *common name*, *formula*, *how obtained* (give equation, if possible), and *two uses* for eight important calcium compounds.

Optional Questions

1. Describe the manufacture of ordinary glass. Write an equation for its formation.
2. Why must we eat foods that contain calcium compounds? List three foods that are rich in calcium.
3. Why cannot rock phosphate (calcium phosphate) be used as a fertilizer? How is it converted to calcium superphosphate (monocalcium phosphate)?
4. Describe the manufacture of Portland cement. Discuss the hardening of cement. What is concrete? Reinforced concrete?
5. Calcium belongs to the group of metals known as the alkaline earths. List other metals belonging to this group.
6. Explain the use of calcium sulfite in paper making.

UNIT 41. HYDROCARBONS AND OTHER ORGANIC COMPOUNDS

Experiments



Apparatus for Preparing Methane

1. Methane. *Demonstration.*

(a) Arrange apparatus as shown in the figure above. Mix a little dry sodium acetate with 3 times its bulk of dry soda-lime [$\text{NaOH} + \text{Ca}(\text{OH})_2$] on a piece of paper. Fill the hard glass generating test tube $\frac{1}{3}$ full with this mixture. Heat cautiously and collect the methane gas in test tubes or collecting bottles. Do not collect the first few bubbles of air.

(b) Half fill one test tube with gas. Allow air to replace the remaining water. Bring the mixture of methane and air to a flame. Result?

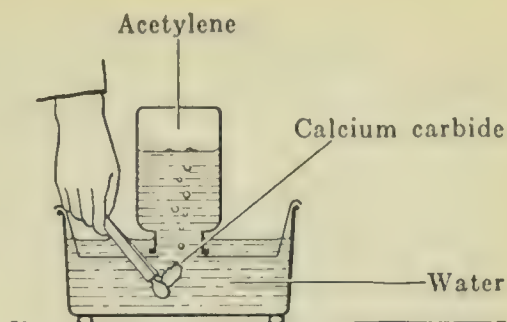
(c) Repeat Exp. 1 (b), trying to obtain varying mixtures of methane and air in the test tube. Results?

(d) To a test tube of methane gas add a few drops of dilute bromine water. Shake. Is there a color change? No color change indicates that the bromine does not combine with the methane.

2. Acetylene.

(a) Invert a bottle filled with water in a pan of water as in the figure at the right. Place a small piece of calcium carbide under the mouth of the bottle. When the bottle is filled with gas, place it mouth upward on the table.

Standing at arm's length, bring



Apparatus for Preparing Acetylene

a flame to the bottle. Result? The gas is acetylene.

(b) Repeat Exp. 2 (a) two or three times. Try to obtain varying mixtures of acetylene gas and air in the bottle. Bring a flame to the bottle in each case. Results?

(c) Test the solution in the pan with litmus paper. Result?

(d) To a test tube of acetylene gas prepared in a similar way add a few drops of dilute bromine water. Shake. Result? Compare with results in Exp. 1 (d) above.

3. Gasoline and kerosene.

Caution: Keep bottles of gasoline and kerosene away from flames.

(a) Dip the end of a clean glass stirring rod in gasoline. Bring the rod to a flame. Result? Repeat, using kerosene. Result?

(b) Place 5 drops of gasoline and 5 drops of kerosene on separate watch glasses. Set aside to evaporate. Note the time needed for evaporation in each case.

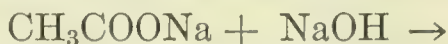
(c) Place 2 or 3 drops of gasoline in a wide mouth bottle. Cork the bottle and shake vigorously. Remove the cork and **standing at arm's length**, bring a burning splint to the mouth of the bottle. Result?

Copy the following in your notebook, supplying the word or words needed to make each a complete statement. Underline the words you supply.

Observations and Questions on Experiments

1. Methane is collected by the —?— method indicating that it is —?— in water.

2. Representing soda-lime, $[\text{NaOH} + \text{Ca}(\text{OH})_2]$, as NaOH , complete the equation for the preparation of methane:



3. Methane gas is (combustible, noncombustible) —?—.

4. Applying a flame to a mixture of methane and air causes —?—.

5. Complete the equation for the burning of methane:



6. Bromine (combines, does not combine) —?— with methane as shown by —?—. This shows that methane is (a saturated, an unsaturated) —?— compound.

7. Acetylene is collected by the —?— method indicating that it is —?— in water.

8. When calcium carbide reacts with water, the resulting mixture changes litmus from —?— to —?— indicating the formation of a —?— in addition to the formation of acetylene gas.

9. Complete the equation: $\text{CaC}_2 + \text{H}_2\text{O} \rightarrow$

10. Acetylene gas is (combustible, noncombustible) —?—.

11. Applying a flame to a mixture of acetylene and air causes —?—.

12. Complete the equation for the complete combustion of acetylene: $\text{C}_2\text{H}_2 + \text{O}_2 \rightarrow$

13. Bromine (combines, does not combine) —?— with acetylene as shown by —?—. Acetylene is a (saturated, unsaturated) —?— hydrocarbon.

14. Gasoline adhering to a glass rod burns —?— while kerosene under similar circumstances burns —?—.

15. A flame applied to a mixture of gasoline vapor and air causes —?—.

16. Assume gasoline to have the formula C_6H_{14} and kerosene to have the formula C_9H_{20} . Complete the following equations for their complete combustion in air:



17. Assuming that kerosene is burned in a limited supply of air, complete the equation: $\text{C}_9\text{H}_{20} + \text{O}_2 \rightarrow$

Conclusions

1. Hydrocarbons are compounds of the two elements —?—.
2. The hydrocarbon methane can be made by the reaction of —?— and —?—.
3. Acetylene can be made by the reaction of —?— and —?—.
4. The complete combustion (with plenty of air) of hydrocarbons produces —?— and —?—. The incomplete combustion (with a limited supply of air) of hydrocarbons produces —?— and —?—.
5. The vapor of a hydrocarbon, as well as any other combustible vapor, mixed with air forms an —?— mixture.

Supplementary Exercises

1. A common name for methane is —?—. It is used as —?—.
2. Methane is the first member of a series of (saturated, unsaturated) —?— hydrocarbons called the —?— series. The general formula for this series is —?—.
3. Ethylene gas (formula: —?—) is a hydrocarbon found in commercial —?— gas. Besides its use as a fuel, it is also used —?—.
4. Ethylene is the first member of a series of (saturated, unsaturated) —?— hydrocarbons called the —?— series. The general formula for this series is —?—.
5. Acetylene gas is used for —?—. It is the first member of the series of (saturated, unsaturated) —?— hydrocarbons called the —?— series. The general formula for this series is —?—.
6. Benzene (formula: —?—) should be distinguished from benzine, the mixture of hydrocarbons obtained from petroleum oil. Benzene is obtained from —?—, and is used for —?—. Benzene is also the first member of the series with the general formula —?—.
7. The hydrocarbons found in petroleum oil belong to the —?— series of hydrocarbons.
8. Five commercial hydrocarbons obtained from petroleum oil in the order in which they boil off in the process of —?— are —?—.

9. In the "cracking process" some of the more complex hydrocarbons of petroleum are —?—.

10. Two important halogen substitution products of methane are —?— (formula: —?—) and —?— (formula: —?—).

11. Gasoline is dangerous for cleaning because —?—.

12. Substances having the same percentage composition (same simple formula) but having different —?— because of the different arrangement of the atoms in the molecule (different —?— formulas) are called —?—.

13. Write the structural formulas for the following: methane, ethane, ethylene, acetylene, benzene, chloroform.

14. Copy and complete the following tables in your notebook:

Alcohols. (General formula is —?—)

NAME	FORMULA	HOW OBTAINED	USES
Methyl alcohol (wood alcohol)			
Ethyl alcohol (alcohol)			
Glyceryl alcohol (glycerine)			
Phenol (carbolic acid)			

Organic Acids. (General formula is —?—)

NAME	FORMULA	HOW OBTAINED	USES
Formic acid			
Acetic acid			
Oxalic acid			
Tartaric acid			

Organic Salts (Esters). (General formula is —?—)

NAME	FORMULA	HOW OBTAINED	USES
Glyceryl nitrate (nitroglycerine)			
Sodium stearate (soap)			
Ethyl acetate			
Amyl acetate (banana oil)			
Glyceryl stearate (stearin or fat)			
Methyl salicylate (oil of wintergreen)			

Carbohydrates. (General formula is —?—)

NAME	FORMULA	HOW OBTAINED	USES
Starch			
Cellulose			
Simple sugar (glucose, dextrose, etc.)			
Complex sugar (sucrose, lactose, maltose, etc.)			

Optional Questions

1. What is the formula of formaldehyde? How is it prepared? What are its uses?
2. What is the formula of ether? How is it prepared? What are its uses?
3. Describe the commercial process used in extracting and refining sugar.

4. Why is concentrated sulfuric acid used in making so many esters?
5. Relate the story of Perkin and his discovery of the first coal tar dye.
6. How is bakelite made? What are some of the properties that make possible its many uses?
7. To what class of organic compounds does acetone belong? What is its formula? Its uses?
8. Write the structural formulas of normal butane and isobutane.
9. Distinguish between aliphatic and aromatic series of organic compounds.
10. Explain the importance of cellulose in three different industries.
11. Write a 300-word essay on the development of synthetic organic chemistry starting with the work of Woehler.

UNIT 42. SOAP

Experiments

1. Place about $\frac{1}{3}$ of a test-tubeful of olive oil (or beef fat or cottonseed oil) in an evaporating dish. Add $\frac{1}{2}$ a test-tubeful of a 30% solution of sodium hydroxide. Add a test-tubeful of alcohol. Heat the dish gently with constant stirring until you no longer can smell the alcohol. **(If the alcohol catches fire, remove the burner and cover the dish with a watch glass, a pan or asbestos sheet.)**

When the mixture has jellied, pour the warm mixture into a beaker half filled with a saturated salt solution. Stir and allow to settle. The soap separates out on top.

Skim off some of the soap with a spatula or glass plate. Put some in a test tube containing water. Shake. Any suds produced? Wash your hands with a little of it.

2. Test some soap solution with litmus paper. Result?

3. Add a few drops of soap solution to some calcium sulfate solution (hard water) in a test tube. Shake. Any suds produced? Compare with the effect of a few drops of soap solution on distilled water.

4. To a test tube half filled with water add a few drops of kerosene. Shake. Allow to stand. Result? Add a little soap solution. Shake. Allow to stand. Result?

5. Put a very small amount of lampblack into each of two test tubes half filled with water. To one test tube add some soap solution. Shake both test tubes. Allow to stand for about 5 minutes. Result? Filter the contents of each test tube. Examine the filtrates. Results?

Copy the following in your notebook, supplying the word or words needed to make each a complete statement. Underline the words you supply.

Observations and Questions on Experiments

1. Alcohol is not essential in the making of soap. In this experiment it is used to —?—, but it is not used commercially because —?—.

2. The two essential ingredients used in making soap are —?—.

3. Complete the equation: $C_3H_5(C_{17}H_{35}COO)_3 + 3NaOH \rightarrow$
(fat)

4. An important byproduct in the making of soap is —?—.

5. The purpose of adding salt to the soap mixture is to —?—.

The process is called —?— the soap.

6. Soap solution changes litmus from —?— to —?— indicating the presence of —?— ions.

7. Ordinary soap is a salt of the base —?— and of some —?— acid.

8. With the usual hard water, soap forms —?— instead of —?—.

9. When kerosene and water are shaken together and allowed to stand, the kerosene —?—.

10. When soap is added to a mixture of kerosene and water and the mixture is shaken, the kerosene forms an —?— with the water. The soap is the —?— agent and forms a film around the small droplets of kerosene oil preventing these droplets from —?—.

11. When a mixture of lampblack and water is filtered, the lampblack (does, does not) —?— go through the filter paper. When soap is added to a mixture of lampblack and water, the lampblack (does, does not) —?— go through the filter paper.

12. Soap has the effect of putting the lampblack in —?— suspension. The soap acts as a protective colloid.

Conclusions

1. Ordinary soap is made by heating —?— with —?—. The process of making soap is called —?—.

2. Soap is separated from the reacting ingredients and its byproduct, namely, —?—, by treating the mixture with —?—. The soap collects on —?—.

3. Soap cleans by —?— grease and thereby loosening the —?— away from the grease to which it had been clinging. The soap also helps get the —?— into —?— suspension.

Supplementary Exercises

1. Wood ashes leached with water produce a solution of —?— which when boiled with —?— produces soap.

2. Ordinary soap, usually sodium stearate ($\text{NaC}_{17}\text{H}_{35}\text{COO}$) is a salt of the (strong, weak) —?— base —?— and of the (strong, weak) —?— acid —?—. Soap will therefore hydrolyze in water solution to give (an acid, an alkaline) —?— reaction. This —?— character of soap partly explains its cleaning action.

3. Ordinary hard soap is a compound of the metal —?—, and ordinary liquid or soft soap is a compound of the metal —?—.

4. Castile soap is made from —?— oil.

5. Laundry soaps usually contain a little free —?—.

6. Floating soaps contain —?—.

7. Scouring soap contains some fine gritty material such as —?— or —?—.

8. Soap powders usually contain —?— and —?—.

9. An important "filler" for soap is —?—.

10. Three basic cleaners used with water other than soap are: —?—.

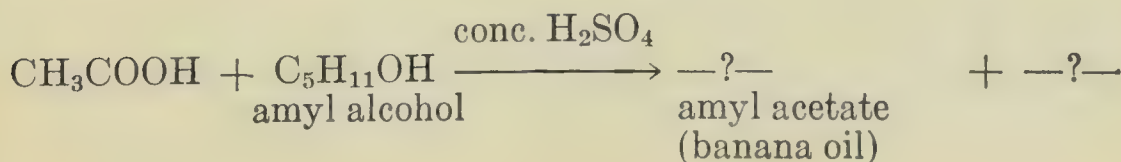
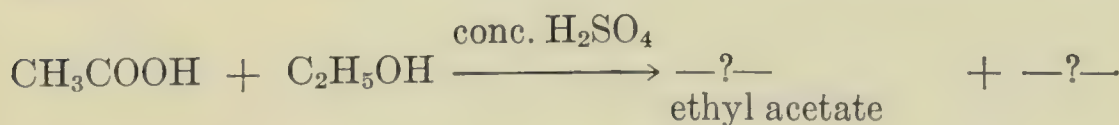
11. Three so-called "dry" cleaners which are not used with water are —?—.

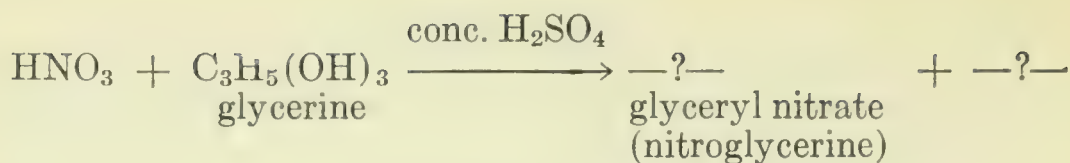
12. All of the common cleaners —?— or —?— grease.

13. Hard water soaps usually contain —?— in addition to soap.

14. Soap is an ester. An ester may be defined as the compound other than water formed when —?— reacts with —?—. Esters are often called —?— salts.

15. Complete the following equations for the preparation of esters:





Optional Questions

1. Why is blueing used in washing clothes? How is it applied?
2. Why is starch used in some clothes? How is it applied?
3. When is dry cleaning preferable to soap and water cleaning?
4. What precautions must be observed in the use of dry cleaners?
5. What do medicated soaps contain?
6. What is tincture of green soap?

UNIT 43. PAINTS AND LACQUERS

Experiments

1. Oil paints.

(a) Dip a wooden splint or wooden applicator stick into each of various kinds of oils and set aside to dry or harden. (Some oils that might be tested are: cottonseed oil, raw linseed oil, boiled linseed oil, motor oil, turpentine, and a mixture of boiled linseed oil and Japan drier.) Examine each sample at the end of laboratory period, on the next day, and on a number of successive days thereafter. Results?

(b) Dip each of two wooden splints into boiled linseed oil. Leave one splint exposed to the air and place the other splint into an air-tight sealed tube. Examine each splint 24 hours later. Results?

(c) Obtain some thick *white lead* and try to paint a small wood surface with it. Then add turpentine to the paint and mix until the paint has the consistency of cream. Then paint again with it. Which covers the surface better?

(d) *DEMONSTRATION*—Prepare a mixture of $\frac{1}{3}$ boiled linseed oil, $\frac{1}{3}$ turpentine, and $\frac{1}{3}$ Japan drier in a shallow dish. Saturate a piece of absorbent cotton with this mixture. Set the cotton aside on a metal or porcelain surface and observe its change in temperature using a thermometer or by touching the outside of the metal or porcelain surface. Results?

(e) Paint 2 small wood surfaces—one with a white *lead* paint, and one with white *zinc* paint. While still moist, subject each surface to either a jet of hydrogen sulfide (H_2S) gas or a drop of ammonium sulfide [$(\text{NH}_4)_2\text{S}$]. Any change in color?

(f) To test paints to see whether they contain lead or zinc bases, shake 3 cc. of the paint with 5 cc. of carbon tetrachloride. Allow to stand a few minutes and then pour off the top liquid. (This process is known as *decanting*.) The carbon tetrachloride will separate most of the linseed oil and turpentine from the

pigments. Some of the bottom residue is then placed in a cavity of a charcoal block and reduced with a blowpipe flame. If lead is present, a silvery globule of lead will form in the cavity. If zinc is present, a yellowish coating will form. After this coating has cooled it becomes white. (Also see Unit 34 for further tests.)

2. Lacquer paints.

A few crystals of a dye (Gentian Violet or Malachite Green or some other dye), 5 cc. of collodion, and 5 cc. of amyl acetate are ground together forming a thorough mixture. This mixture is further diluted with equal volumes of collodion and amyl acetate until the consistency of thin cream is obtained. A drop of olive oil may be added. The mixture is then ready to be sprayed, brushed, or poured or used in other ways to coat surfaces such as wood or glass or even paper. Individuals in the class may try various ways of painting with this *lacquer* paint. Various dyes can be used.

Copy the following in your notebook, supplying the word or words needed to make each a complete statement. Underline the words you supply.

Observations and Questions on Experiments

1. The oils tested which show signs of hardening or "drying" are —?— and —?—.

2. The oils tested which remain liquid and which do not harden or "dry" are —?— and —?—.

3. In order for an oil such as linseed oil to harden it must come in contact with —?—.

4. Linseed oil is an unsaturated compound. This means its molecules can readily —?—.

5. The tough elastic solid produced when linseed oil "dries" is caused by the absorption and combination of the linseed oil with —?—.

6. The kind of linseed oil which hardens or "dries" most effectively is the —?—.

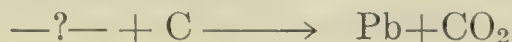
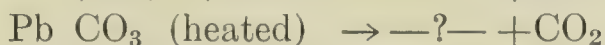
7. The substance used as a "thinner" by means of which the paint can be spread more easily, evenly, and thinly is —?—.

8. Linseed oil can be made to “dry” faster by the addition of —?—.

9. When linseed oil “dries” the temperature —?— proving that the chemical process of —?— is taking place.

10. Hydrogen sulfide or other sulfur compounds will cause white “lead” paints to —?— while white “zinc” paints are —?—.

11. Complete the following equations:



12. Amyl acetate (banana oil) serves the purpose of —?— in a lacquer paint.

13. Lacquer paints dry quickly because —?—.

14. Collodion is a solution of nitrocellulose in a mixture of the solvents —?— and —?—. These solvents evaporate quickly.

Conclusions

1. The “drying” of linseed oil is in reality a process of —?—.

2. Paints contain two essential ingredients; a PAINT BASE which has the characteristics of —?—; and a VEHICLE which when mixed with the paint enables —?—.

3. An example of a paint base is —?—; of a paint vehicle is —?—.

4. Paints generally have added to them also a substance called a “thinner” (e.g. —?—) which serves the purpose of —?—.

5. Paints generally have added to them a substance called a —?— which gives color to the paint.

6. Lacquer paints dry by the rapid evaporation of a solvent such as —?— leaving behind —?—.

Supplementary Exercises

1. The most commonly used paint “vehicle,” linseed oil, is obtained from the —?— plant.

2. Boiled linseed oil is made by heating raw linseed oil with —?—.

3. Substances which are put into paint to speed up the oxidation or "drying" of paint are known as —?—. An illustration of such a substance is —?—.

4. Tung oil (or Chinese wood oil) is being used extensively today as a —?— in many paints and enamels.

5. Oil varnish differs from paint mainly in containing no —?—. Spirit varnish contains resin or a similar solid dissolved in a volatile solvent such as —?—.

6. Some well known paint bases, all white in color, are:

White lead (composition: —?—)

Zinc white (composition: —?—)

Lithopone (composition: —?—)

Titanox (composition: —?—)

7. Pigments added to paint are generally different colored metallic —?—, —?—, and —?—.

8. Dyes are finding increasing use as paint pigments though they generally have the disadvantage of —?—.

9. The most common and effective black pigments are —?—, the amorphous carbon made by the partial combustion of hydrocarbon bases such as methane; and —?— made by a similar burning of hydrocarbon oils such as kerosene.

10. Some substances used as pigments are: for yellow, —?— and —?—; for red, —?— and —?—; for blue, —?—.

11. Some characteristics of a good paint base are: —?—, —?— and —?—.

12. Cheap materials added to paint to reduce the cost are called —?—. They are often referred to as adulterants. However they often serve a beneficial purpose by helping to —?— and —?—.

13. Some substances commonly used for the purpose indicated in Ex. 12 are —?—, —?—, —?— and —?—.

14. Michael Angelo created a new form of painting when he incorporated pigments in wet plaster on walls for decorative purposes. Such paintings are known as —?—.

15. A "skin" coating forms on the surface of an open can of paint because —?—.

Optional Questions

1. Discuss the advantages and disadvantages of a "lead" paint; of a "zinc" paint.
2. Explain why paint is used in our homes.
3. Explain why paint pigments should be insoluble.
4. Explain the use of "red lead" in painting steel or iron structures.
5. Distinguish between a paint and a varnish; between a paint and a stain.
6. What is lithopone?
7. What does the word "extender" mean in connection with paint? What purpose does an extender serve?
8. What is meant by the term "unsaturated oil"?
9. In mixing paints for various color effects, why are yellow, red and blue paints or pigments considered essential?
10. What is shellac? How is it prepared?
11. What is an enamel? Is it the same as the enamelware used for kitchen utensils? Explain.
12. Discuss the problem of paint for boats.
13. Discuss the use of rubber latex in paints.
14. What is furfural? Discuss its use in the paint industry.

UNIT 44. FOOD INGREDIENTS

Experiments

1. Starch.

(a) Put a pinch of starch into a test tube half filled with water. Shake thoroughly. Add two or three drops of iodine solution. Color? Boil until the color disappears. Then cool the test tube by holding it under running cold water. Results?

(b) Test freshly cut pieces of potato, bread, and apple with drops of iodine solution. Results?

2. Sugar.

(a) To 5 cc. of Benedict's (or Fehling's) solution in a test tube add a few drops of glucose solution. Boil for 2 minutes. The red precipitate formed is an indication of the presence of a simple sugar such as glucose or fructose.

(b) Repeat Exp. 2 (a) using a solution of cane sugar (sucrose). Is the red precipitate obtained? To 5 cc. of some sucrose solution in a test tube add a few drops of dilute hydrochloric acid. Heat to boiling. Cool. Slowly add powdered sodium carbonate until it is slightly alkaline (test with litmus). Now test again with Benedict's solution. Result?

(c) See Unit 29 to review the effect of concentrated sulfuric acid on carbohydrates such as starch, cellulose, and sugars.

3. Fats and oils.

(a) Rub some fat or place a drop of vegetable oil such as olive oil on a piece of paper. Hold to the light. Result?

(b) Put a little of a food (ground corn or peanuts) in a test tube. Add a few drops of gasoline or benzol. Shake vigorously. Allow to stand. Pour 2 or 3 drops of the clear top liquid onto a piece of paper. Allow the solvent (the gasoline or benzol) to evaporate. Hold to the light. Result?

4. Proteins.

Place a little egg albumen (the white of a hard boiled egg) into a test tube. Add 1 or 2 drops of concentrated nitric acid. Color? Wash with water. Add 2 or 3 drops of concentrated

ammonium hydroxide. Color? These color changes are used in testing for protein matter in foods.

5. Mineral matter.

Burn a *small* sample of food (meat, cheese, or bread) on a porcelain crucible cover until all the black carbon is oxidized away. What is left?

Copy the following in your notebook, supplying the word or words needed to make each a complete statement. Underline the words you supply.

Conclusions

1. Test for starch. When —?— is added to cold starch, a —?— color appears. Upon heating, this color —?—. Upon further cooling, the color —?—. The formula for starch is —?—.

2. Test for a simple sugar. Add some —?— solution to a little of the unknown food and heat. A —?— shows the presence of a simple sugar. Three examples of simple sugars are: —?—. The formula for a simple sugar is —?—.

3. Cane sugar, beet sugar, and milk sugar (formula —?—) do not reduce Benedict's (or Fehling's) solution as do simple sugars. They can be converted to the simple sugars by boiling in water acidulated with —?—. This process is called —?— of sugar.

4. When concentrated sulfuric acid is added to sugar, starch, or cellulose (carbohydrates), the effect is —?— because —?— is removed.

5. Test for fats or oils. A fat or oil can be dissolved out of food by a solvent such as —?—. The solution is placed on a piece of paper, the solvent is —?—, and the presence of the fat or oil is shown by the —?— effect on the paper.

6. Test for proteins. Add a drop or two of —?— to a little of the unknown food. A —?— color which turns to an —?— color upon application of —?— shows the presence of a protein in the food.

7. Test for mineral matter. Burn some of the unknown food

until all the black carbon is burned away. A —?— shows the presence of mineral matter.

Supplementary Exercises

1. Carbohydrates are compounds of the three elements: —?—, in which the two last named elements are in the ratio of —?— to —?—.
2. Three examples of foods rich in carbohydrates are: —?—.
3. The function of carbohydrate foods is to —?—.
4. Fats and oils are complex compounds of the three elements: —?—.
5. Three examples of foods rich in fats or oils are: —?—.
6. The function of fats and oils as food is to —?—.
7. Proteins are complex organic compounds. They are the chief food ingredients that contain the element —?—.
8. Three examples of foods rich in proteins are: —?—.
9. The function of protein foods is to —?—.
10. Mineral matter found in foods contains compounds of some of the following eight elements: —?—.
11. The function of mineral matter in food is to —?—.
12. Three foods rich in mineral matter are: —?—.
13. Vitamins are complex substances found in minute quantities in many of our foods. Their presence in our diets prevents diseases such as the following four diseases: —?—.
14. Four foods rich in vitamin content are: —?—.
15. Milk is considered a perfect food because —?—. Milk is pasteurized by heating it to —?— degrees C. for about —?— minutes and then cooling it rapidly and keeping it cold. Milk is so treated in order to —?—.
16. The energy or heat value of food is measured in terms of —?—.
17. A balanced diet means —?—.

Optional Questions

1. What are the usual legal requirements for the composition of milk that may be offered for sale?

2. What is certified milk? modified milk? sterilized milk? powdered milk? homogenized milk?
3. How is butter made? What is oleomargarine?
4. Explain why the appetite may not be an accurate guide to the food needs of the body.
5. Explain why many foods must be cooked before being eaten.
6. Explain what happens to food in the process of digestion.
7. What is a food adulterant? Name some adulterants in common use.
8. What is a food preservative? Name some preservatives in common use.
9. What are hormones? Cite some illustrations. Explain why hormones are important.
10. Write a 300-word essay on the history and importance of vitamins.

UNIT 45. CONSUMER TESTS OF CERTAIN FOODS

NOTE: *Usually adulterants in foods are cheaper materials than the food, which are added to increase the volume of the product thereby lowering its quality. When such adulterants are used, the consumer pays more for the food because its quality has been misrepresented. Usually adulterants of this kind are not harmful to health. Materials classed as adulterants include also certain harmful preservatives.*

NOTE: *In making the various tests listed below, it is necessary to run control tests. Foods to which have been added a small quantity of the adulterant are tested at the same time that the suspected food is tested.*

Experiments

1. Canned foods.

A. Note the appearance of the can. If it is bulged or otherwise out of shape, this may be due to the putrefaction and fermentation of the contents. Putrefaction and fermentation are accompanied by the formation of gas. To detect such food, cover the end of the can with water and puncture the can with 2 or 3 small holes. If bubbles of gas come through the water, the food is unfit for use.

B. *Preservatives* are generally harmful, especially if used in large amounts. The following tests apply to bottled foods and meats as well as to canned foods.

1. *Sulfites.* To about an ounce of the food material (if necessary add water to make a thin liquid) add a few small pieces of zinc and about 10 cc. of dilute hydrochloric acid. Allow to stand in a warm place or over a very low flame for a few minutes. If sulfites are present, hydrogen sulfide gas with its characteristic odor of rotten eggs will be given off.

2. *Other preservatives.* Mash about 2 tablespoons of the food, add 100 cc. of water, stir, and filter through cloth. The filtrate is tested as follows:

(a) *Borax or boric acid.*

Collect 5 cc. of the filtrate in a watch glass. Add 4 drops of concentrated hydrochloric acid. Stir. If borax or boric acid is present, turmeric paper dipped in this solution will turn cherry red when dried over a radiator or other warm place. A drop of ammonium hydroxide on this paper, when cold, will turn the cherry red color to greenish black.

(b) *Sodium benzoate, benzoic acid, salicylic acid, and saccharin.*

To 20 cc. of the filtrate in a separatory funnel, add 10 drops of concentrated hydrochloric acid. Mix. Add 30 cc. of carbon tetrachloride. Stir gently, allowing to stand, and separate the bottom carbon tetrachloride layer. Divide into 2 parts.

(1) Place one part of the carbon tetrachloride layer in a watch glass and allow the carbon tetrachloride to evaporate. The residue, when tested, will be very sweet if *saccharin* is present.

(2) Sodium benzoate or benzoic acid are shown by the presence of small flat crystals in the residue. If either is present, on warming the watch glass the odor of gum bezoin should be observed.

(3) To the other half of the carbon tetrachloride layer add 5 cc. of water and a very small crystal of a ferric salt such as ferric alum. Shake in a separatory funnel and allow to separate. Salicylic acid will cause the upper layer or the line separating the upper layer from the lower to be purple in color.

C. Coal-tar dyes. (Often found in tomato catsup, jams, and jelly.) Use about a tablespoonful of the food digested with about 150 cc. of water. Mix well and filter through a cloth. Add 3 drops of hydrochloric acid and heat gently in a beaker. Put a piece of white woolen cloth or a piece of cheesecloth into the mixture and stir and heat for about 10 minutes. Remove the cloth and rinse in clear water. If the cloth has become colored, dyes are present in the food. To confirm this test, dissolve the aniline color out of the cloth with about 100 cc. of a dilute solution of ammonium hydroxide. (Natural fruit color on the cloth will not be affected.) Neutralize this ammonium hydroxide solution, and then make it slightly acid with dilute hydrochloric acid (use litmus as an indi-

cator.) Then stir a fresh piece of cloth in this acid solution. If this second piece of cloth is dyed and retains its color when rinsed in water, aniline dyes are shown to be present.

D. *Starch.*

Dissolve some of the washed food in a little hot water, shake or stir thoroughly and filter. To the cold filtrate add a few drops of tincture of iodine. An intense dark blue coloration indicates the presence of starch.

II. *Meats.*

Consumers should buy government inspected meats. The government inspection stamp is an indication that the meat is from a healthy animal and that it was handled under sanitary conditions. A good grade of meat should be free from spotty discolorations and unusual odors.

The usual preservatives found in meat products are *sodium sulfite* (which produces a reddish color in imitation of freshly cut meat), *borax*, and *boric acid*. These preservatives are tested for as indicated under *Canned foods* above. For the borax and boric acid tests, the meat must be finely divided and warmed with water to extract the juices. The common adulterant in sausages is *starch* which can be tested for as indicated above.

III. *Milk.*

The adulterants sometimes added are *water*, *coloring matter*, and *preservatives*.

A. A sensitive hydrometer (lactometer) provides confirmatory evidence when milk has been diluted with water. The specific gravity of milk should be between 1.027 and 1.033. A hydrometer reading below 1.027 indicates either that the milk is very rich in cream *or* that it probably has been watered. The appearance of the milk is an indication of whether desirable cream or undesirable water is in excess.

When water is used as an adulterant the milk becomes bluish-white and a yellow dye may be added to restore the color. To test for this yellow coal-tar dye, add 25 cc. of concentrated hydro-

chloric acid to a 25 cc. portion of the milk, and heat. A pink coloration indicates the presence of the dye.

B. Annatto, a yellow coloring material, may be detected as follows: Add a little washing soda (sodium carbonate) to the milk, stir, and allow to stand about 3 minutes. Filter and wash the milk off the filter paper. Annatto will color the paper yellow.

C. *Preservatives.* If preservatives are added to milk, the milk will not turn sour in the usual time under ordinary conditions and putrefaction and disease bacteria may accumulate without the usual warning (curdled sour milk).

Formaldehyde (formalin) is a cheap and harmful preservative. It is detected as follows: To 10 cc. of the milk in a test tube add 10 cc. of concentrated hydrochloric acid and shake vigorously. Add 3 drops of a 10% solution of ferric chloride. Stir or shake vigorously. Heat very gently by moving the test tube through a Bunsen flame a few times. Stir or shake vigorously. Allow to stand 5 to 10 minutes. A purple color appears. This color is intensified and darkened by subsequent gentle heating.

NOTE: This test can serve to detect formaldehyde in ice cream. The presence of starch in the ice cream will cause a lighter shade of purple. Starch can be tested as shown in Part I.

NOTE: For the control test add 3 drops of a 40% solution of formaldehyde to 50 cc. of milk.

IV. Butter.

Oleomargarine is a substitute for butter. It is made from mixtures of varying amounts of oleo oil, cottonseed oil, lard, milk, and butter. Although nutritious and a good food, it is cheaper than butter and it is a misrepresentation to sell it as butter.

Renovated butter is made by heat treating rancid butter to drive off the rancid odors.

Fresh butter can be distinguished from renovated butter and oleomargarine as follows: Heat a small lump of the food in an iron spoon or combustion spoon over a gas flame. Fresh butter melts quietly forming a lot of foam with many small bubbles. Oleomargarine or renovated butter sputters explosively and produces very little foam.

V. *Coffee.*

A. Drop some finely ground coffee into a beaker containing cold water. Real coffee will float and will not discolor the water for a few minutes. Most of the adulterants commonly used will sink to the bottom and cause an immediate brown discoloration of the water.

B. To test for starchy adulterants often used in coffee, take some freshly brewed coffee, let it cool, dilute it with water until it is not too strongly colored, and then test with tincture of iodine. Starch will produce a characteristic dark blue color.

VI. *Baking Powders* (see Unit No. 33).

The volume of carbon dioxide liberated from equal weights of various baking powders may be used as a measure of their relative effectiveness and also of the amount of adulterant added.

Weigh out 3 grams of the baking powder and place it in a small vial. Have ready an 8 oz. bottle filled with a saturated solution of sodium chloride (400 g. per liter) inverted in a trough filled with the same saturated solution (brine). Insert the vial containing the baking powder into this bottle. Carbon dioxide will be liberated displacing some of the brine in the flask. The amount of displaced brine is a measure of the carbon dioxide content of the baking powder.

VII. *Bleached Flour.*

To some of the flour in a bottle, add enough colorless gasoline to cover it. Unbleached flour will color the gasoline yellow. Bleached flour will not affect the color of the gasoline.

VIII. *Eggs.*

The freshness of an egg is indicated by the following tests:

A. Hold the egg in front of a candle flame or other bright light. The center should look uniformly rose colored and clear. Numerous dark spots show that the egg is not fresh.

B. A fresh egg placed in a dish of cold water will sink.

C. A fresh egg generally has a somewhat rough shell.

IX. Fish.

If the fish is fresh, the flesh should be firm, and the eyes and gills should be bright.

X. Poultry.

A fresh chicken should have soft feet, a smooth skin, and a soft cartilage at the end of the breast bone.

Questions on Experiments

1. What is meant by the term *adulterant*? Why should consumers object to the adulteration of foods?
2. What is meant by the term *preservative*?
3. Describe a simple way of detecting the putrefaction or fermentation of canned foods.
4. Describe the chemical tests employed in the detection of borax or boric acid in foods.
5. What is saccharin? Is it a food? Why is its use for sweetening most prepared or canned foods prohibited by law? How may its presence in food be detected?
6. Describe the chemical tests used in the detection of starch. Is starch harmful as an adulterant in food? Cite a case in which its use as an adulterant is considered objectionable.
7. What effect does the addition of sodium sulfite have on meat? How may the presence of this adulterant be detected?
8. What is the most common adulterant added to milk? What are the legal requirements, in your home city or state, regarding the composition of milk?
9. How can you test for the presence of the preservative formaldehyde in milk or ice-cream?
10. What differences are there between butter, renovated butter, and oleomargarine?
11. How could you show that coffee is adulterated with starch?
12. How could you show that flour has been bleached?
13. Baking powders liberate what gas as a leavening agent? When is this gas liberated from the baking powder?
14. How could you show whether an egg is fresh?

15. Why must we eat food? What are the functions of food in our bodies?

16. What are the essential food ingredients? Mention the most important function of each ingredient.

17. Name some diseases that occur if vitamins are absent from food.

18. Name the known vitamins and, for each vitamin, name a disease which may occur if that vitamin is not present in the diet.

19. Name some foods rich in vitamins and the vitamins they contain.

20. Has the function of vitamins in food been over-stressed in recent advertising? Discuss both sides of this question.

21. Cite instances of misleading advertisements of food products. Explain why you believe them to be misleading.

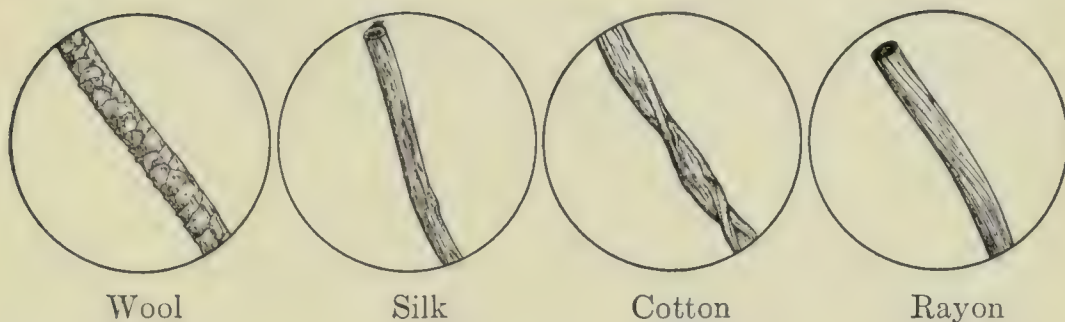
22. Name some food adulterants in common use. Point out disadvantages in buying food containing these adulterants.

23. Name some preservatives used commonly in foods. Point out possible advantages and disadvantages of these preservatives.

24. Write a report on the most recent Food and Drug Act passed by the Congress. How does it protect the consumer?

UNIT 46. TEXTILES

Experiments



Textile Fibers as Seen under the Microscope

1. Examine fibers of cotton, wool, silk, linen, and rayon under the high power microscope. The instructor will have the slides prepared and set up. Make drawings of each textile fiber to show differences in appearance as seen under the microscope.

2. Burn small pieces of each of the raw textiles listed in the table on page 188. Make a similar table in your notebook and record the speed of burning, the kind of flame, the odor in each case, and the nature of any ash produced.

3. Place a small strip of each of the 5 textiles listed in the table in 5 separate test tubes. Add 5 cc. of a 20% solution of sodium hydroxide. Boil for 5 minutes. Record your results in your table.

4. Try the dissolving effect of warm concentrated hydrochloric acid on each of the textiles. Record results in the table in your notebook.

5. Place small pieces of linen and cotton (equal size) in separate test tubes. Add 5 cc. of concentrated sulfuric acid to each. After 2 minutes remove (with a stirring rod) each textile from the acid. Rinse with water and then with dilute ammonium hydroxide. Results? (The linen will not be affected as much as the cotton because of the presence of silicon compounds in the flax fiber which are insoluble in ordinary reagents.)

TABLE

TEXTLE	SPEED OF BURNING	NATURE OF FLAME	ODOR	ASH	EFFECT OF NaOH SOLUTION	EFFECT OF CONCENTRATED HCl	EFFECT OF CONCENTRATED H ₂ SO ₄
Cotton							
Wool							×
Silk							×
Linen							
Rayon							×

Copy the following in your notebook, supplying the word or words needed to make each a complete statement. Underline the words you supply.

Supplementary Exercises

1. A mineral fiber used for fireproof materials is —?—.
2. Wool is (a vegetable, an animal) —?— fiber.
3. Silk is (a vegetable, an animal) —?— fiber.
4. Cotton is (a vegetable, an animal) —?— fiber.
5. Linen is (a vegetable, an animal) —?— fiber.
6. Rayon is a —?— for silk. Most of the rayon manufactured today is made by the —?— process. Cellulose from —?— or —?— is treated with —?— solution. The soda cellulose thus formed is then treated with —?— and water to form a syrupy liquid called viscose. This viscose is then pumped through fine platinum nozzles and hardened into fine filaments in a bath of —?— and other chemicals. These filaments are further purified and treated to make rayon yarn.
7. Rayon is inferior to silk in —?—. It is superior to silk in —?—.
8. Rayon looks like the textile fiber —?— but burns like the fiber —?—.
9. Mercerized cotton is cotton treated with —?— to give it a —?— luster and additional strength.

10. Woolen garments should not be left on a hot radiator because —?—.

11. Linen is preferable to cotton for handkerchiefs because —?—.

12. Precipitating certain inorganic salts such as tin phosphate in the fibers of silk is called —?— the silk.

Optional Questions

1. What is cellophane? How is it made? How does it compare with rayon as to composition and method of manufacture?

2. Is an asbestos curtain a textile? Explain.

3. Why is woolen clothing warmer than cotton clothing?

4. Explain how the natural silk fiber is made.

5. Discuss the Chardonnet process of making a substitute for silk.

UNIT 47. REMOVAL OF SPOTS AND STAINS

Experiments

NOTE: *The instructor should have prepared various small pieces of cotton, linen, wool, and silk each stained with one of the stains listed in the table on the opposite page. Each student should be given a few pieces of stained cloth and asked to remove the spots.*

PROCEDURE:

If you cannot recognize the spot or stain, use carbon tetrachloride or soap and water, or experiment with some of the substances listed in the table. If you recognize the nature of the spot or stain, use the solvent or cleaning agent recommended in the table.

In applying a cleaning fluid such as carbon tetrachloride, place a clean piece of white blotting paper or cloth under the stain. Apply the cleaning fluid to the spot or stain with a piece of cloth and rub from the outside towards the center to avoid the formation of a "ring."

Keep in mind that alkalies should not be used with wool or silk. Also keep in mind that the color of a colored piece of cloth will be affected by bleaching agents such as sodium hypochlorite (NaOCl) and hydrogen peroxide (H_2O_2). In case of doubt always test some inconspicuous piece of the garment with the reagent you expect to apply.

Consult the table on the opposite page when cleaning the pieces of cloth. Make a similar table in your notebook.

TABLE OF SOLVENTS OR CLEANSING AGENTS

	STAIN	COTTON OR LINEN	WOOL OR SILK
1	Grease	Carbon tetrachloride <i>or</i> gasoline.	
2	Ink	When fresh, water, then soap and water, <i>or</i> warm solution of ammonium oxalate, then water, <i>or</i> NaOCl, then oxalic acid, then water.	When fresh, water, then soap and water, <i>or</i> warm solution of ammonium oxalate, <i>or</i> warm milk.
3	Paint	Xylol, <i>or</i> turpentine, <i>or</i> carbon tetrachloride.	
4	Tar	Benzol, <i>or</i> soap and turpentine, <i>or</i> carbon tetrachloride.	
5	Fruit	Hot water, then ammonium hydroxide, <i>or</i> oxalic acid.	Hot water, then ammonium hydroxide, <i>or</i> tincture of soap.
6	Grass	Alcohol, <i>or</i> tincture of soap.	
7	Iron rust	Dilute hydrochloric acid, <i>or</i> oxalic acid, <i>or</i> lemon juice.	Vinegar, <i>or</i> lemon juice.
8	Molasses	Water.	
9	Chewing gum	Ether, <i>or</i> toluene.	
10	Blood	Water, then soap and water, <i>or</i> NaOCl solution if necessary.	Water, then soap and water, <i>or</i> alcohol if necessary.

Copy the following in your notebook, supplying the word or words needed to make each a complete statement. Underline the words you supply.

Supplementary Exercises

1. The common basic cleansers are dissolved in —?— when used.
2. Four of the common basic cleansers are —?—.
3. A dry cleaner means one which does not use —?—.
4. Three substances used as dry cleaners are: —?—.
5. The cleansers or stain removers usually work by —?— the grease or the specific substance causing the stain.

6. The safest dry cleaning agent is —?— because —?—.
7. Javelle water is the common name for —?— whose formula is —?—. It is used as a —?— agent.
8. Bleaching powder is the common name for —?— whose formula is —?—.
9. Washing soda should not be used in cleaning woolens because —?—.
10. A grease stain is more easily removed from a woolen suit than from a silk dress because —?—.
11. Corroded or tarnished metals may be cleaned by means of abrasives such as —?— and —?— and by means of acids such as —?— and —?—.
12. Aluminum should not be cleaned with alkalies such as —?— because —?—. Aluminum kitchen utensils are usually cleaned with —?— as an abrasive, aided by a mild vegetable soap.
13. Silverware can be conveniently cleaned by boiling it in an —?— dish containing a solution of —?—.

Optional Questions

1. Why do writing ink stains gradually turn brown?
2. Account for the formation of rust stains while washing clothes.
3. Explain the cleaning action of soap.
4. Explain the use of sodium hypochlorite solution as a cleaner.

UNIT 48. PHOTOGRAPHY

Experiments

NOTE: *These experiments should be performed in a darkened laboratory.*

1. Chemistry of photography.

(a) To each of three *clean* test tubes (two of them covered with dark or black paper to prevent the penetration of light) containing 5 cc. of silver nitrate solution, add about 5 cc. of potassium bromide solution. Examine the precipitate of silver bromide in the exposed test tube.

(b) Permit the exposed test tube to stand in direct sunlight (or in some powerful artificial white light) for 2 minutes. Does the precipitate (AgBr) change color? Now add 5 cc. of the developer solution. Result?

(c) To one of the unexposed test tubes add 5 cc. of the developer solution. Remove the black paper. Does the developer solution change the precipitate (AgBr) which was not exposed to light?

(d) To the other unexposed test tube add 10 cc. of hypo solution. Shake the tube thoroughly. Remove the black paper. What happens to the precipitate of silver bromide?

NOTE: *Photographic films or plates and sensitized print paper (Velox paper) are coated with a gelatin emulsion carrying particles of silver bromide and silver iodide. Making a print (or positive) on a piece of Velox paper involves the same procedure as making a negative from a photographic film or plate. The first step is exposure of the silver compounds to light. The second step is to continue the chemical action started by the light by using a developer. The third step is to dissolve out the silver compounds that were unexposed to light by using hypo. The final product has a black deposit of silver wherever there was a penetration of light.*

2. Printing a picture.

If the school is equipped with a photographic "dark room"

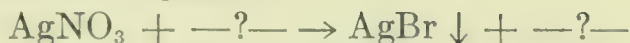
the instructor should demonstrate to small groups of students how to print a picture on a piece of Velox paper from a negative. After the demonstration students can practice individually. (If the "dark room" is not available try to darken the laboratory or some other room as effectively as possible.)

Using red or amber light fix up the printing frame with the dull side of the negative against the glossy side of the print paper. Send light, for about 8 seconds, from a 100-watt electric lamp held about 10 inches away, through the negative and onto the print paper. (The time of exposure depends on the negative, the kind of paper, and on the amount of light.) Examine the paper in red light. Any noticeable effect? Then place the exposed print paper in the developer solution until the picture comes out in clear outline. (If the picture comes out too quickly and looks too dark, the print was exposed too long. If the pictures comes out too slowly, the print was underexposed.) Rinse the print paper with water and then put it into the hypo solution for at least 10 minutes. Then rinse well with water and allow to dry.

Copy the following in your notebook, supplying the word or words needed to make each a complete statement. Underline the words you supply.

Observations and Questions on Experiments

1. Complete the equation:



2. Before exposure to light, silver bromide has a ---?--- color.
3. After exposure to sunlight for about 2 minutes, the color of silver bromide ---?---.
4. Upon the addition of the developer solution, the silver bromide exposed to light changes in color to ---?--- because of the production of ---?--- colored ---?---. ($\text{AgBr} \rightarrow \text{Ag} + \text{Br}$)
5. The developer solution (changes, does not change) ---?--- the silver bromide that was not exposed to light.
6. The hypo solution ---?--- the silver bromide that was not exposed to light.

7. Photographic films, plates, and sensitized papers can be safely handled in red or amber light because —?—.

8. In taking a picture with a camera, the light from the —?— portions of the scene or object goes through the opened shutter of the camera and starts the decomposition of the —?— on the film or plate. After developing, those parts of the film or plate acted on by the light turn —?—. Thus the —?— parts of the scene or object appear —?— on the negative. In the fixing process, the —?— dissolves the —?— not affected by light. That part of the negative now appears transparent.

9. In printing a picture from a negative, —?— light is sent through the negative on to the —?— paper. The light does not go through the —?— portions of the negative, where there is a deposit of —?— colored —?—. The transparent part of the negative, indicating the —?— portions of the original scene or object, permit light to go through to the —?— paper. Then when this paper is developed and fixed, the —?— colored deposit of —?— will appear wherever the original object appeared dark colored or black.

Conclusions

1. When a light sensitive film or piece of paper (coated with an emulsion of —?— and gelatin) is exposed for a short time to —?— light, the light —?—.

2. The purpose of the developing solution (containing a mild organic (oxidizing, reducing) —?— agent) is to —?—.

3. The purpose of the fixing solution (containing the substance —?—) is to —?—.

4. The bright or light portions of the objects or scenes photographed are represented on the negative by —?—. Light and dark portions of the object photographed appear in (true, opposite) —?— fashion on the negative.

5. In making a print from a negative, the process is essentially the same as —?—. On the print, light and dark portions of the object photographed appear in (true, opposite) —?— fashion.

Supplementary Exercises

1. A silver compound, other than silver bromide, used in photography is —?—.
2. The kind of light rays most effective in decomposing silver compounds are the —?— rays obtained from —?—, —?— lamps, and —?— lamps.
3. The kind of light rays that have practically no effect on silver compounds are the —?— rays.
4. An example of a substance used in developers is —?—.
5. Toning a print means replacing the silver by either —?— or —?—.
6. A sepia print is produced by converting the silver to —?—.
7. Motion picture film and other films are often made of cellulose acetate instead of the usual cellulose nitrate because —?—.
8. Two chemical reactions other than the ones in photography that are caused by light are: —?—.
9. Two chemical reactions that result in the evolution of light are: —?—.

Optional Questions

1. How does X-ray photography differ from ordinary photography?
2. Describe the operation and use of a photoelectric cell.
3. Write a 200-word essay on the history of photography.
4. Mention three important changes, other than those in photography, caused by sunlight.
5. Describe how a blueprint is made. Explain the chemistry involved. Why do builders use blueprints rather than photographs?
6. Write a 300-word essay on a modern natural color-process in photography.
7. Write a 200-word essay on the spectroscope, including its make-up, its uses, and its value.
8. What connection is there between sunlight and vitamins?
9. Compare light waves with x-rays and radio waves.
10. What is a quantum? A photon?

UNIT 49. COLLOIDS

NOTE: Construct a beam of light by means of a bright lamp and lens arrangement or by means of a mirror and sunlight. This beam is to be used in experimenting with the various solutions described below. A stereopticon or similar projector could be used. If the solutions are in beakers, look from above at the bottom of the beaker to see if the beam is visible.

Experiments

I. Formation of colloids.

A. Peptization—breaking down larger particles into smaller particles of colloidal size.

1. Place a small quantity of coarse starch in a test tube. Fill the test tube with water and shake. Filter through filter paper. Test the filtrate with a few drops of dilute tincture of iodine. Did the starch go through the filter paper?

2. Repeat part 1, but this time use a small quantity of starch which was ground thoroughly in a mortar with a pestle. Results?

3. Repeat part 1, but this time heat the starch in the water until the water is about to boil. Results?

4. Try the beam of light through distilled water in a clean beaker. Is the path of light visible? Repeat using a filtered dilute sugar solution. Is the path of light visible? Repeat using some of the solution from part 3. Results?

B. Coagulation—combining particles of molecular size into larger groups of colloidal size.

1. Make a saturated solution of sulfur in alcohol (or isopropyl alcohol). Filter. Add a little of this filtrate, drop by drop, to about 200 cc. of distilled water in a clean beaker until the presence of colloidal sulfur is shown by a faint opalescence. Now add more of this solution of sulfur in alcohol until a precipitate forms. What happened? Filter through filter paper. Examine the filtrate. Test the filtrate with a beam of light as in part A above. Results?

2. Fill about two-thirds of a test tube with a freshly filtered

concentrated solution of sodium thiosulfate (hypo) in water. Add 5 cc. of dilute hydrochloric acid. Filter. Test the filtrate with a beam of light as above. Results?

3. Fill a clean beaker half full of distilled water. Bring the water to a boil and then add a 10 per cent solution of freshly filtered ferric chloride, drop by drop, until a deep red coloration forms. Filter and test the filtrate with a beam of light as above. Results? What colloid was formed?

C. *Emulsions.*

Mix about 10 cc. of water and 5 cc. of kerosene. Shake thoroughly. Allow to stand. What happens? Add a few drops of liquid soap and shake thoroughly again. Allow to stand. Result?

II. Uncharged colloidal suspension (emulsoid colloid).

Mix 3 grams of *Jello* or a similar gelatin preparation with 12 cc. of boiling water in a test tube. Shake. Stir and allow to stand for 24 hours. Result? Explain.

III. Protective colloids.

1. Add a little lampblack to each of two test tubes half full of water. To one test tube add a few drops of liquid soap. Shake both test tubes thoroughly. Allow to stand. How do they compare? Filter each. Results?

2. To 15 cc. of sodium chloride solution add some silver nitrate solution. Result? To 15 cc. of a sodium chloride solution add an equal volume of a 1 per cent gelatin solution and then an equal volume of silver nitrate solution. Result? What is the effect of the gelatin?

IV. Adsorption Phenomena.

1. Repeat Experiments 2 and 3 of Unit 30.

2. Make a dilute solution of methylene blue (or methyl violet) in a test tube two-thirds full of water. Add this solution to a test tube containing about 1 inch of powdered charcoal (activated charcoal is desirable if available). Shake thoroughly and filter. Note color of filtrate. Repeat, using copper sulfate instead of the dye. Results?

3. Rinse a test tube which has contained a water solution of

methyl violet (or a similar dye) with distilled water about 6 or 7 times. Then to the same test tube, which seems to be free of dye, add some alcohol (or iso-propyl alcohol) and shake. Does the color appear again? Explain.

Copy the following in your notebook, supplying the word or words needed to make each a complete statement. Underline the words you supply.

Observations and Questions on Experiments

1. When a mixture of water and coarse particles of starch are filtered, the starch (does, does not) —?— pass through the filter paper.

2. When a mixture of water and finely ground particles of starch are filtered, some of the starch —?— pass through the filter paper. The starch in this case is said to be in —?— solution.

3. Starch heated in water forms a —?— solution.

4. The beam of light is (visible, invisible) —?— in distilled water; —?— in a mixture of starch and hot water; —?— in a sugar solution; —?— in the filtrate from finely ground starch and water.

5. When a solution of sulfur in alcohol is dropped into water, a —?— solution of sulfur in —?— is produced. This is indicated by the appearance of the solution and by the beam of light appearing —?—.

6. Hydrochloric acid added to sodium thiosulfate (hypo) produced a —?— solution of —?— in water.

7. Ferric chloride solution added to hot water produces a colloidal solution of —?— in water.

8. A good emulsifying agent for water and kerosene is the substance —?—.

9. Emulsoid colloids such as gelatin or gelatin dessert preparations when mixed with hot water and allowed to stand produce the effect of —?—. This is because the gelatin —?—.

10. Lampblack will settle more readily in water (with, without) —?— soap.

11. Lampblack can form a colloidal solution in water if a substance such as —?— is added. This substance is known as a —?— colloid.

12. Sodium chloride and silver nitrate solutions, when mixed, cause a precipitation of —?—. When gelatin is added previously, these solutions when mixed will produce —?—.

13. Powdered or amorphous charcoal or carbon can remove colloidal odors and colors by the process called —?—.

14. Amorphous or activated charcoal cannot remove the color from a copper sulfate solution because —?—.

15. Methyl violet or a similar dye seemingly washed out of a test tube with water, can impart a color to some alcohol added to the test tube because —?—.

Conclusions

1. Starch particles can be broken down to colloidal size by —?— and —?— the starch.

2. Peptization is the process of —?—.

3. The path of a beam of light is visible if seen through a —?— solution.

4. By the process of coagulation, particles of colloidal size are made from —?—.

5. When two ordinary non-miscible liquids are thoroughly mixed with the aid of a protective colloid, they form a mixture called —?—.

6. The swelling of certain colloids such as gelatin is caused by the colloid taking up —?—.

7. Substances added to colloidal solutions to prevent the colloids from coagulating or precipitating out are called —?—.

8. Colloids show the property of —?— when, because of their large surface exposure, they can cling to other surfaces.

Supplementary Exercises

1. Colloidal particles are intermediate in size between the very small particles in —?— and the comparatively large particles of substances in —?—.

2. The visibility of a beam of light through a colloidal solution is known as the —?— effect.

3. Particles in colloidal solution will readily go through a filtering device such as —?— but will not go through a filtering device such as —?—.

4. The special filtration process for the separation of colloidal particles is known as —?—.

5. Particles in the —?— are electrically charged as whole particles.

6. Particles of ferric hydroxide, gelatin, albumin and hemoglobin are charged —?—, and particles of starch, silver chloride, oil emulsions, and clay are charged —?—.

7. The movement of electrically charged colloidal particles under the influence of an electric current is known as —?—.

8. The peculiar zig-zagging movement of colloidal particles as seen under a microscope is known as the —?— movement. This is explained as the movement caused by —?— of the surrounding liquid hitting the —?— particles.

9. The noticeable beam of sunlight passing through dust particles in a room is an illustration of the —?— effect.

10. The man usually considered to be the “Father of Colloidal Chemistry” is —?—.

11. A machine which will grind coarse material into very fine material of colloid size is known as a —?—.

12. Some examples of “sols” (also called “lyophobic” and “suspensoids”) are —?—, —?—, and —?—.

13. Some examples of “gels” (also called “lyophiles” and “emulsoids”) are —?—, —?—, and —?—.

14. Some examples of protective colloids are the —?— in mayonnaise, the —?— in milk, and the —?— in ice cream.

15. The precipitation of colloids by neutralizing the charge on the colloidal particles in smoke has been used in the —?— process in the chimneys and flues of many factories.

16. The —?— of milk coagulates forming curds when the milk turns sour (or an acid is added).

17. Colloidal particles have been shown to be as large as —?— mm. and as small as —?— mm.

18. The special kind of microscope used in studying colloidal particles is known as an —?—.

19. Acheson used —?— as a protective colloid in keeping graphite in colloidal suspension in water.

20. The property of adsorption is largely due to —?— of surface exposed by colloidal particles.

21. In the tanning of leather —?— is used to coagulate the gelatin in the hide.

22. In photography, the colloidal silver salts in film, plates, and sensitized papers are protected by —?—.

23. In the digestive processes in our bodies, colloid foods such as starch and proteins must be changed so that they can pass through —?—.

24. —?— colloids such as gelatin absorb more or less of the dispersion medium.

Optional Questions

1. Explain how soap cleanses by adsorption.
2. Explain the difference between absorption and adsorption.
3. What is the composition of the usual rubber cement?
4. Why is human milk considered superior to cow's milk for infants?
5. What is the Cottrell process for the elimination of the smoke nuisance?
6. Explain the formation of river deltas from the point of view of colloid chemistry.
7. Explain why the following statement is not accurate:
"Matter can be classified as colloidal or crystalloidal."
8. What is syneresis?
9. What is meant by a reversible colloid?
10. State the properties characteristic of matter in the colloidal state.

UNIT 50. CONSUMER TESTS OF COMMONLY USED MATERIALS

1. Dentifrices.

Tooth pastes and tooth powders help somewhat in removing food particles from the teeth. They may also make the removing process somewhat more pleasant. The consumer, however, should beware of dentifrices that claim to remove "film" from teeth. If this film is not food residue, it is generally harmful to remove it by using abrasive pastes or powders. Dentifrices which are too abrasive may wear away the enamel and expose the soft part of the teeth thereby hastening decay.

Also beware of extravagant claims that by merely using a certain toothpaste or powder one can cure pyorrhea or whiten the teeth or prevent decay, or cure halitosis or disinfect the mouth. A dentifrice which is safe for you to use daily can not do any of these things.

The following ingredients found in dentifrices may be harmful or undesirable if the dentifrice is used daily or oftener: acids, potassium chlorate, borax, sugar, starch, and harsh abrasives such as pumice.

Experiments

(a) The best and lowest priced way to help your toothbrush do its work is to dip it into a warm solution of either salt (NaCl) or baking soda (NaHCO_3) or a mixture of both before you brush your teeth. Either solution can be made easily by dissolving about $\frac{1}{4}$ of a teaspoon of the solid in one-half glass of warm water. Another inexpensive and safe dentifrice is ordinary Precipitated Chalk (CaCO_3) which can be purchased at drug-stores for about 35 cents a pound. Oil of wintergreen or oil of peppermint can be added (about 15 drops per lb.) if flavoring is desired.

(b) To test for pumice or similar gritty material in a tooth powder or toothpaste.

(1) To about $\frac{1}{3}$ of a teaspoon of the tooth powder in a test

tube, slowly add $\frac{1}{3}$ of a test tube of dilute hydrochloric acid. Shake thoroughly. Filter. Wash the residue with water by pouring a test tube of water over it and allowing the water to drain. If the residue then feels gritty when rolled between your fingers, it probably contains an abrasive such as sand or some silicate found in pumice.

(2) Obtain a new, clean microscope slide, free from scratches. Place on it about $\frac{1}{2}$ inch of toothpaste (or about $\frac{1}{4}$ teaspoon of tooth powder made into a paste with a few drops of water). Hold a smooth well-worn nickel coin perpendicular to the slide, and in a direction parallel to the long edges of the slide, rub with the curved edge of the coin. About 200 *light* strokes (100 double strokes) should be used. Avoid excess pressure. Remove the toothpaste by holding the slide under running hot water. Examine in a strong light to see if the slide was scratched. For confirmation the test should be repeated. A control test also should be run with the same coin and another portion of the same slide using only water on the slide.

(c) To test for starch.

Add two or three drops of a dilute iodine solution (iodine crystals, 0.3 gm.; potassium iodide, 1.5 gm.; water, 100 cc.) to a small quantity of the dentifrice and mix thoroughly. Starch is indicated by the familiar blue-black coloration.

(d) To test for borax. See Unit 45, part I.

(e) To test for potassium chlorate. Test for the potassium. See Unit 36. To test for the chlorate, add a few drops of concentrated hydrochloric acid to a small portion of the toothpaste or tooth powder. Either dentifrice will turn yellow due to dissolved chlorine gas and other substances which are formed. A piece of litmus paper placed in this mixture would be bleached.

(f) To test for sodium perborate.

Sodium perborate should be used only upon the advice of a doctor or dentist as its excessive use may cause injury to the mucous membranes of the mouth.

Sodium perborate, when dissolved in water, undergoes chemical action and oxygen gas is evolved slowly.

Test for the sodium. See Unit 36. Turmeric paper, if moistened with a water solution of sodium perborate and acidulated with hydrochloric acid, becomes brown in color when the paper is dried. Moisten the dried paper with ammonium hydroxide. A greenish-black color is produced.

2. Cosmetics.

Cosmetics are substances applied to the skin or hair, designed to improve their appearance. Whether they really make people look better is unimportant to us. The use of cosmetics is so widespread and so many millions of dollars are spent on these "beautifiers" that they warrant some consideration and study in a chemistry course.

A. Face powders.

Face powders usually contain a few finely powdered ingredients. These are generally zinc oxide, talc (magnesium silicate), zinc stearate, precipitated chalk (calcium carbonate), perfume, and coloring matter. The coloring material is usually carmine from the cochineal insect *or* yellow ochre from oxides of iron *or* various aniline dyes from coal tar. Starch, usually from rice, is considered an objectionable ingredient.

B. Rouge and lipstick.

Rouges are simply highly colored face powders. The coloring materials are the ones mentioned above, usually carmine or aniline dyes. When rouge is made into cakes, a substance such as gum tragacanth (or starch) is used as a binding agent. Petrolatum is the usual base for a paste rouge.

Lipsticks are generally made by mixing the same kind of coloring materials used in face powders and rouges with wax, cocoa butter, or pomade (perfumed lard or fat or fatty material).

Experiments

1. To test for starch. See Experiments under Dentifrices.

In a test tube, boil a little of the face powder in about five times its volume of water for about one minute. Allow to cool.

Add a drop of iodine solution. A blue-black or purple color indicates the presence of starch.

2. To test for a zinc compound. See Unit 43.

To a little face powder in a test tube add twice its volume of dilute hydrochloric acid. Add ammonium hydroxide to neutralize. Then bubble hydrogen sulfide gas through the solution (or add ammonium sulfide solution). A white precipitate (ZnS) indicates the presence of zinc.

Also try the cobalt nitrate test for zinc described in Unit 36.

3. To test for a carbonate, talc, calcium, and iron.

(a) To test for a carbonate. To a little face powder in a test tube, add twice its volume of dilute hydrochloric acid and boil for about three minutes. While the acid is being added, effervescence and the evolution of a gas probably indicates the presence of a *carbonate*.

(b) To test for talc. If the powder does not dissolve entirely, it probably contains *talc* (magnesium silicate).

(c) To test for calcium. To another portion of the liquid, add ammonium hydroxide to neutralize. Then add a few drops of ammonium oxalate solution. A white precipitate (calcium oxalate) indicates the presence of *calcium*.

(d) To test for an iron compound. To another portion of the liquid, add a drop of nitric acid and a few drops of a solution of potassium sulfocyanate (KSCN). A bright red coloration indicates the presence of an *iron* compound.

4. Composition of a Typical Face Powder.

1. Talc	50 parts by weight
2. Koalin	20 parts by weight
3. Zinc Oxide	15 parts by weight
4. Zinc Stearate	10 parts by weight
5. Precipitated Chalk	5 parts by weight
6. Coloring Matter	Small amount to suit
7. Perfume	Small amount to suit

C. Face creams.

Face creams are generally cold creams or vanishing creams.

Cold creams contain greasy or fatty substances which make the skin softer. Vanishing creams usually are greaseless and serve substantially the same purpose as soap.

Some of the fatty substances in cold creams are: animal or vegetable fats such as stearin; lard; olive oil; lanolin, a fatty material obtained from sheep's wool; petroleum jelly (vaseline); or paraffin, obtained from petroleum. Petroleum products do not turn rancid as do the animal and vegetable fats or oils.

Discount claims that certain face creams are skin foods. The skin is fed from within, not from without.

Some objectionable ingredients sometimes found in cold creams are compounds of lead, bismuth and mercury. These are harmful both to the skin and the body.

Experiments

1. To test for a *Lead*, *Bismuth*, or *Mercury* compound.

Place a tablespoon of the cold cream in an evaporating dish or beaker and add 10 cc. of concentrated nitric acid. Heat for about 3 minutes, add 10 cc. of water and filter. Divide filtrate into 3 portions.

(a) To test for a mercury compound. To one portion of the filtrate, add a few drops of stannous chloride solution. A grayish precipitate shows the presence of a *mercury* compound.

(b) To test for a mercury compound. To another portion of the filtrate add a brightly polished strip of copper. A silvery white deposit of mercury will confirm the presence of a mercury compound.

(c) To test for a lead or bismuth compound. To another portion of the filtrate, add a few drops of potassium chromate solution. A yellow precipitate indicates either *lead* chromate or *bismuth* chromate. The lead chromate is soluble in sodium hydroxide solution while the bismuth chromate is not.

2. To make a cold cream.

Mix the following ingredients in a beaker or deep dish. Mix thoroughly and heat until liquid. Heat very gently.

White Mineral Oil	45%
White Beeswax	13%
Spermaceti	6%
Then, while stirring,	
Powdered Borax	1%
Hot water (200°F.)	35%

NOTE: *Perfume is added and stirred in while the preparation is cooling. The mixture should be put into jars while still warm.*

D. Hair tonics, dyes, bleaches, and depilatories.

Hair tonics. Most hair tonics are of little or no value. Massaging the scalp tends to keep it in a healthy condition with or without the use of hair tonic.

Hair dyes. "The usual safe hair dye is not very effective and the effective hair dyes are not very safe." Hair dyes are especially dangerous to the eyes or to an open cut or abrasion in the scalp. The most effective type of hair dye is said to contain paraphenylene-di-amine, or a similar organic compound. However, under recent Federal laws, dyes of this type are required to be indicated on the label as follows:

"Caution. This product contains ingredients which may cause skin irritation on certain individuals, and a preliminary test according to accompanying directions should first be made. This product must not be used for dyeing the eyelashes or eyebrows; to do so may cause blindness."

The metallic hair dyes are very effective, but are inclined to be dangerous to the health of many people. These dyes contain usually a lead or silver salt which, in the presence of air and light, produces a dark colored or black deposit of either the metal or its sulfide. Sometimes these dyes contain a dangerous copper salt mixture.

The vegetable hair dyes are the safest. Generally they are less effective and less permanent than metallic dyes. The vegetable coloring agent most commonly used is henna which produces reddish shades. Constant use of henna may cause the hair to

become brittle and coarse. Henna is sometimes adulterated with harmful metallic salts of copper, lead or silver. Hair "restorers" are really hair dyes.

Hair bleaches. The most commonly used hair bleach is a solution of hydrogen peroxide. The solution usually sold for both disinfecting and bleaching purposes is a 3% solution. This is generally safe to use, but may after long usage make the hair brittle and dry. A powder type bleach which is safe to use is a mixture of magnesium (or calcium) carbonate and hydrogen peroxide solution ("white henna"). A type which is harmful to the scalp and hair is one which contains sodium perborate.

Depilatories. Depilatories or hair removers are substances which remove hair by dissolving the protein of which it is composed. However, these same substances usually will dissolve or irritate the skin also and are therefore dangerous to use. The most effective way of removing unwanted hair is by the electric needle process in the hands of a skilled operator. The usual depilatories contain starch mixed with a sulfide of either sodium, barium, or calcium. Any of these sulfides dissolves hair and other proteins.

Experiments

1. To test for silver salts. Add a few drops of dilute hydrochloric acid to some of the filtered liquid hair dye. A white precipitate which is soluble in ammonium hydroxide indicates the presence of a *silver* salt.

2. To test for lead salts. Add a few drops of potassium chromate solution to some of the filtered liquid hair dye. A yellow precipitate which is soluble in sodium hydroxide indicates the presence of a *lead* salt.

3. To test for copper salts. Heat some of the filtered liquid hair dye on the end of a clean platinum or nichrome wire. A characteristic green coloration indicates the presence of a *copper* salt.

Ammonium hydroxide added in excess to a solution containing a copper salt produces a deep blue coloration.

4. To test for hydrogen peroxide. Soak a strip of filter paper in a dilute solution of lead acetate. Then dip this paper into a

fresh solution of ammonium sulfide or hydrogen sulfide. Brownish-black lead sulfide is formed. Then place this blackened strip of paper in some of the hair bleach solution. If the black color disappears, the presence of hydrogen peroxide is indicated.

5. To test for magnesium or calcium carbonate. Add some dilute hydrochloric acid to a small quantity of the bleach. Effervescence and the evolution of a gas (carbon dioxide) which turns limewater milky is an indication of the presence of a *carbonate*.

6. To test for sodium perborate. See Experiment (f) under Dentifrices above.

7. To see how a hair dye or "restorer" works. Soak some white or gray wool threads (wool is a protein material similar to hair) in a lead acetate solution. Then dip the threads in a sodium sulfide (or other soluble sulfide) solution. Rinse the wool and dry. Note the brownish-black deposit of lead sulfide.

8. To see how a depilatory works. Place 2 or 3 hairs in a dish and cover with a paste composed of barium sulfide (*or* sodium sulfide *or* calcium sulfide) and water. Note that the hairs are practically dissolved in a few minutes.

D. Deodorants.

Deodorants are usually substances which are essentially astringents which reduce the size of the pores in the skin, and thereby reduce the quantity of perspiration given off. Some well known deodorants are solutions of alum (potassium aluminum sulfate), aluminum chloride, and zinc chloride. One deodorant is an approximately 20% solution of aluminum chloride in water.

If you touch a little of any of these solutions to your tongue, you will notice its astringent, or puckering, effect. Such substances may cause severe irritation to sensitive skin. Generally they are exorbitantly high in price.

Deodorant powders and pastes contain one or more of a number of the ingredients indicated in the make-up of a typical deodorant powder listed at the top of page 225.

Experiments

1. Composition of a deodorant powder.

Powdered talc -----	18 parts by weight
Zinc oleate -----	8 parts by weight
Boric acid -----	5 parts by weight
Magnesium carbonate -----	4 parts by weight
Sodium perborate -----	3 parts by weight
Sodium bicarbonate -----	2 parts by weight

Each ingredient is a powder. The ingredients are mixed thoroughly and perfumed.

2. The various ingredients of deodorant solutions or powders can be identified as indicated in the Experiments listed for the cosmetics described earlier in this unit or by reference to Unit 45.

Observations and Questions on Experiments

1. What is a dentifrice? What is its function?
2. Why are dentifrices usually mildly alkaline?
3. Why should we object to pumice or similar gritty material in a tooth powder or toothpaste?
4. Describe a simple way of detecting such gritty material.
5. What objections are there to the use of starch in a dentifrice? How would you identify starch by a chemical test?
6. Why should borax not be used in a dentifrice?
7. What purpose does sodium perborate serve in a dentifrice?
8. By a chemical test, how could you demonstrate that a dentifrice contains sodium perborate?
9. What are cosmetics? Give reasons for and against their use.
10. Name some ingredients found in face powders. Why is starch generally considered an objectionable ingredient?
11. By a chemical test, how could you show the presence of a carbonate, such as precipitated chalk, in either a dentifrice or a face powder?
12. Name some objectionable ingredients that might be found in a face cream.
13. Is there such a thing as a "skin food" that can be applied

externally? Explain.

14. How could you identify the presence of a lead compound in a cosmetic?

15. Name some ingredients found in cold creams.

16. Is the use of a hair tonic desirable? Explain.

17. Discuss the statement: "The usual safe hair dye is not very effective, and the effective hair dyes are not very safe."

18. How could you show the presence of a silver compound in a hair dye?

19. Indicate the composition of the most common hair bleaches.

20. What are depilatories? What chemical substances are the active ingredients of most depilatories?

21. What objections are there to the use of chemical depilatories?

22. Name some common ingredients of deodorant powders.

23. What is the composition of a typical deodorant solution? How do these solutions reduce perspiration?

24. Discuss the functions of the human skin. What are the sebaceous glands? the sweat glands? hair follicles?

25. The skin of a person may be "allergic" to a particular cosmetic preparation. What does "allergic" mean?

26. About how much is spent annually on cosmetics in the United States? Are we getting our money's worth?

27. Explain how the Federal Food and Drug Act of 1938 tries to help the consumer of drugs and cosmetics.

UNIT 51. PLANTS AND SOILS

A. Plants. (NOTE: See Experiments in Units 44, 45, 50.)

Ten elements have been found essential to the growth of plants. These elements are found in various chemically combined forms (compounds) in plants and in soils. The average plant obtains about 9% of its weight from the air, 89% from water, and 1.5% from minerals in the soil.

The elements essential to the growth of plants and the source from which plants obtain them are:

ELE- MENT	CAR- BON	HY- DROGEN	OXYGEN	NI- TROGEN	PHOS- PHORUS	SUL- FUR	CAL- CIUM	POTAS- SIUM	MAG- NESIUM	IRON
Source	air	and	water and air	From the Soil (Leguminous plants also get Nitrogen from Air)						

Experiments

1. Test for *water* in a plant.

Dry a small pan and weigh it carefully. Fill it with finely cut leaves and stems from the plant. Weigh the filled pan. Subject the filled pan to a temperature of 100°-105° C. for about 5 hours. Cool, weigh, and repeat drying procedure until a fairly constant weight is obtained. Calculate the per cent of water in the plant. It should range from 75 to 95%. Save the dry material. Copy the table below in your note book and complete it.

Weight of pan and plant (before drying).....	grams
Weight of pan.....	"
Weight of plant.....	"
Weight of pan and plant (after drying).....	"
Weight of dried plant.....	"
Weight of the water in the plant.....	"
% of moisture in the plant.....	%

2. Composition of the dry matter of the plant.

a. Testing for *hydrogen, oxygen, and carbon*. Fill a crucible nearly to the top with the dried plant material obtained in part one and heat with a Bunsen or Meker burner flame until the plant

material ignites. Hold a cold dry object (a flask with cold water inside will serve) over the burning material. The droplets of moisture condensing on the cold surface indicate that the remainder of the combined hydrogen in the plant is being oxidized by the oxygen in the air. When the material in the crucible is completely charred, stop heating and observe. The charring effect proves the presence of *carbon* in the plant. Now heat very strongly until only ash remains. The carbon present is removed as carbon dioxide gas ($C + O_2 \rightarrow CO_2$) and some of the *sulfur* present is removed as sulfur dioxide gas ($S + O_2 \rightarrow SO_2$). Save this ash to test for the remaining elements indicated in the chart on page 227.

b. Testing the plant ash. (1) To a small portion (0.5 gram) of the plant ash obtained above, add 5 cc. of distilled water, 5 cc. of concentrated hydrochloric acid, and a few drops of concentrated nitric acid. Heat the mixture to boiling and evaporate *nearly* to dryness. Add 10 cc. of distilled water and stir well. Insoluble compounds present are compounds of *silicon*.

(2) Filter the solution. Wash the residue with distilled water and add washings to the filtrate. Neutralize the filtrate with ammonium hydroxide and add a slight excess of ammonium hydroxide solution. Heat to boiling. A precipitate appears. Filter. The filtrate is saved for (b) below.

(a) To the residue on the filter paper add a few drops of dilute hydrochloric acid and dissolve it. Filter. Catch the liquid that passes through in a test tube. Add a drop of potassium sulfocyanate. A red color proves the presence of an *iron* compound.

(b) Heat the filtrate from part (a) to boiling and add 5 cc. of ammonium oxalate solution. A milky white precipitate shows that a *calcium* compound is present. Filter and divide the filtrate (with washings) into 2 portions.

(1') To one portion of the filtrate from part (b), add slowly (drop by drop) 5 cc. of sodium phosphate (Na_3PO_4) solution and 5 cc. of concentrated ammonium hydroxide. A white precipitate that forms after some time (on standing) proves the presence of a *magnesium* compound.

(2') Evaporate the second portion of the filtrate from part (b)

almost to dryness and then heat strongly until white vapors are no longer given off. Cool and add 2 drops of concentrated hydrochloric acid to the residue. Dip a *clean* platinum or nichrome wire into the residue and then into a colorless Bunsen flame. A bright, persistent yellow flame proves the presence of a *sodium* compound. Observe this flame through a blue cobalt glass. A violet color proves the presence of a *potassium* compound.

(3) To a fresh portion of plant ash (0.5 gram) add 5 cc. of distilled water and 5 cc. of concentrated nitric acid. Heat to boiling, add 10 cc. of distilled water and filter. Divide the filtrate into 3 parts.

(a) To one portion of filtrate from (3), add 2 cc. of silver nitrate solution. A white precipitate of silver chloride (AgCl) proves the presence of a *chlorine* compound. This precipitate is insoluble in nitric acid and soluble in ammonium hydroxide.

(b) To the second portion of the filtrate from (3), add 2 cc. of barium chloride solution. A white precipitate shows the presence of a sulfate and therefore an original *sulfur* compound.

(c) To the third portion of the filtrate from (3), add 5 cc. of ammonium molybdate solution and heat to about 40°C . A yellow precipitate, on standing, shows the presence of a phosphate and therefore an original *phosphorus* compound in the plant.

3. Testing for *nitrogen* compounds in plants.

To a small portion (1 gm.) of dried plant material obtained in part one, add 10 grams of soda lime, $\text{NaOH} + \text{Ca}(\text{OH})_2$. Place in a Pyrex test tube and heat strongly. The evolution of ammonia gas, indicated by its odor and its turning a piece of moist litmus paper held over the test tube from red to blue, proves the presence of *nitrogen* compounds in the plant.

4. Testing for *starch* ($\text{C}_6\text{H}_{10}\text{O}_5$) in plants.

A few drops of dilute iodine solution (0.3 gm. iodine crystals, 1.5 gm. potassium iodide, 100 cc. water) applied to a few slices of potato, corn, or bean shows the presence of starch by the formation of a blue-black or purplish color. To show the presence of starch in a green leaf, it is necessary to kill the leaf by placing it in boiling water for about 3 minutes. Then the leaf is

placed in alcohol and warmed. (*It is dangerous to heat alcohol over an open flame. An electric hot plate is preferable to a Bunsen flame.*) The alcohol dissolves the green chlorophyll from the leaf. Remove the leaf (which is now white). Allow to cool. Iodine solution placed on the leaf will turn it a blue-black wherever starch is present.

B. Soils. (NOTE: *For a detailed procedure of Soil Analysis (which is beyond the scope of this book), the student should obtain a copy of Circular No. 139 (Revised Nov. 1939) U. S. Department of Agriculture, Washington, D. C.—entitled “Method and Procedure of Soil Analysis.” This booklet can be obtained from the Superintendent of Documents, Washington, D. C., for 5 cents.*)

1. Acidity of soil.

Few valuable plants can thrive in a soil which is strongly acid. Unfortunately, because of weathering of the soil, or removal of soil constituents by crops, the basic constituents are removed more quickly than the acid constituents. The loss of basic material is sometimes compensated for by the addition of lime or limestone or similar substances to the soil.

The important metallic or basic constituents are compounds of Na, K, Mg, Ca, Fe, Al, and NH_4 . The important nonmetals which form acid radicals or acid constituents are C, Cl, S, Si, P, and N.

We refer to the degree of acidity of any substance as its pH. A substance which is neutral (neither acid nor basic) has a pH of 7. A pH less than 7 indicates an acid condition. The lower the number, the greater the acidity. Thus a solution with a pH of 5 is more acid than a solution with a pH of 6. A pH more than 7 indicates a basic, or alkaline, condition. The higher the number, the greater the basicity.

Ordinarily, most plants, soils, or nutrient solutions should not have a pH of less than 5. In general, a satisfactory pH range is between 5 and 6.5

Experiments

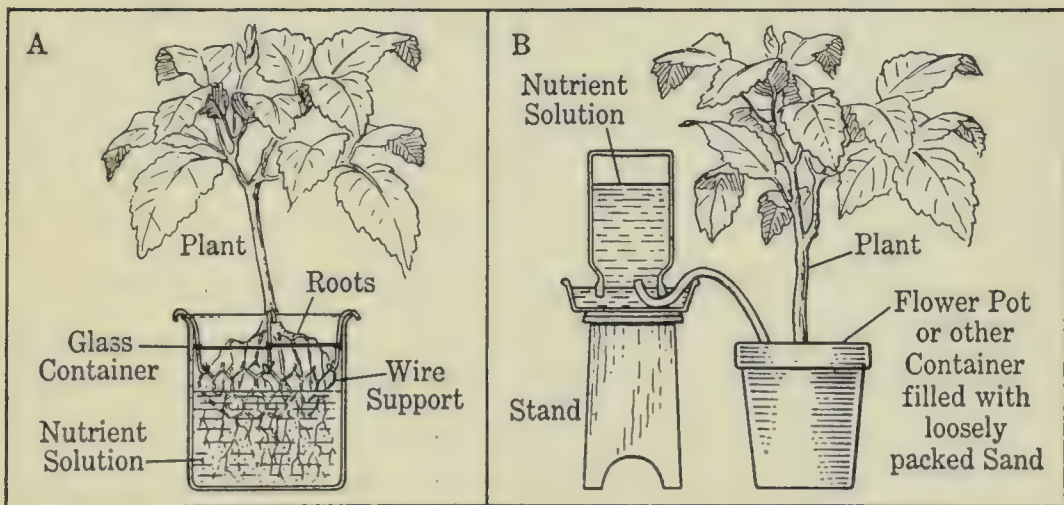
1. Select a representative sample of the soil to be tested. Place some soil on each of two watch glasses, one containing a

strip of blue litmus paper and the other a strip of red litmus paper. Moisten with distilled water to form a thick pasty coating. Press down firmly. After at least 30 minutes, remove the papers from the soils, wash with distilled water, and dry the papers. Results?

2. This experiment can also be performed using Nitrazine paper. The pH value of the soil can thus be obtained through a wider range. At the pH of 4.5, the Nitrazine paper is yellow. As the pH value rises, it turns to green, becomes blue at a pH of 7.5 and violet at a pH of more than 8. A standard color chart can be compared with the Nitrazine paper used to estimate pH fairly accurately. This standard chart may be obtained from the chemical supply company that sells the Nitrazine paper.

2. Growing plants without soil.

The essential constituents of soil necessary for plant growth are indicated clearly in the composition of the nutrient solutions used in various methods for growing plants without real soil.



The following is one of the formulas recommended by the N. J. Agricultural Experiment Station and by Ellis and Swaney, (Ellis and Swaney, *Soilless Growth of Plants*, Reinhold Pub. Co., New York). This book is recommended for those who want a detailed and practical treatment on how to grow plants without soil. Another book recommended is Gericke: *Soilless Gardening*, Prentice Hall, Inc.

	MONO- POTASSIUM PHOSPHATE	CALCIUM NITRATE	MAGNESIUM SULFATE	AMMONIUM SULFATE
Grams per 5 gallons of solution	KH_2PO_4	$\text{Ca}(\text{NO}_3)_2 \cdot 4\text{H}_2\text{O}$	$\text{MgSO}_4 \cdot 7\text{H}_2\text{O}$	$(\text{NH}_4)_2\text{SO}_4$
	5.9 grams	20.1 grams	10.7 grams	1.8 grams

Each of the above chemicals is dissolved in about a quart of water. Then the solutions are mixed and water is added to make five gallons of solution. To each 5 gallons of solution, 10 cc. of a solution containing boron, manganese, and zinc compounds are added to provide the necessary amounts of these elements. This latter solution is made by dissolving 0.8 grams each of boric acid, manganese sulfate and zinc sulfate in one pint of water. A tiny crystal of copper sulfate may be added to this solution. Just before the main solution is to be used, 5 cc. of a solution of either ferrous sulfate or ferric chloride (0.8 gram in one pint of water) is added to each quart of nutrient solution.

Plant roots need air. Air may be bubbled through the nutrient solution in the "A" set-up, or the wire support and the plant may be lifted out of the solution periodically in order to air the roots. Also, as the nutrient solution is added, it is allowed to drip through air thereby carrying dissolved air along with it.

Observations and Questions on Experiments

1. Name 10 elements that have been shown to be essential constituents for the growth of plants. Name 3 that are apt to be deficient in soils after years of crop growth.

2. About what percent of the weight of a plant is obtained from the minerals in the soil? from water? from the air?

3. Describe the procedure by which the percent of water in a plant can be determined.

4. How could you demonstrate by chemical means that a plant contains nitrogen compounds?

5. Describe the procedure by which the presence of starch in a green leaf is demonstrated.

6. What is meant by "nitrogen fixation"? What plants can

“fix” their own nitrogen?

7. Name some chemicals used as fertilizers by farmers.

8. Explain why organic waste materials are used as fertilizers.

9. If a soil is exactly neutral, what is its pH value? What pH value of soil is generally acceptable for plant growth? Is this acceptable soil slightly acid, exactly neutral, or slightly alkaline?

10. If a soil is too acid, what should a farmer add to the soil to neutralize the excess acid?

11. Name the essential chemicals used in making a nutrient solution for soilless gardening or farming. Name the elements which are added in relatively small amounts to the nutrient solution.

12. Explain why plant roots, either in ordinary soil planting or soilless culture, need air.

13. Describe the procedure used in estimating the acidity of a soil.

14. Explain how rock phosphate is treated to make its phosphate content available for use as a fertilizer.

15. Write a report on the various commercial processes used in making nitrate fertilizers from the nitrogen of air.

UNIT 52. REVIEW OF EQUATIONS

Write complete balanced equations for the following reactions.

1. Laboratory preparation of oxygen from potassium chlorate.
2. Heating of mercuric oxide.
3. Burning of magnesium.
4. Burning of sulfur.
5. Burning of carbon.
6. Laboratory preparation of hydrogen by the reaction of a metal and an acid.
7. Electrolysis of water.
8. Action of sodium on water.
9. Burning of hydrogen.
10. Reduction of copper oxide by hydrogen.
11. Laboratory preparation of chlorine.
12. Electrolysis of brine.
13. Chlorine + water.
14. Antimony + chlorine.
15. Hydrogen + chlorine.
16. Laboratory preparation of hydrogen chloride.
17. Zinc + hydrochloric acid.
18. Calcium hydroxide + hydrochloric acid.
19. Sodium chloride + silver nitrate (test for a chloride).
20. Basic anhydride + water as in the reaction between:
 - (a) Calcium oxide + water.
 - (b) Sodium oxide + water.
21. Acid anhydride + water as in the reaction between:
 - (a) Sulfur dioxide + water.
 - (b) Sulfur trioxide + water.
22. Zinc + sulfur.
23. Laboratory preparation of hydrogen sulfide.
24. Incomplete combustion of hydrogen sulfide.
25. Complete combustion of hydrogen sulfide.
26. Zinc chloride + hydrogen sulfide.
27. Copper sulfate + hydrogen sulfide.
28. Laboratory preparation of sulfur dioxide.

29. Sulfurous acid + hydrogen peroxide.
30. Equations for the contact process for making sulfuric acid.
31. Zinc + dilute sulfuric acid.
32. Copper + concentrated sulfuric acid.
33. Potassium hydroxide + sulfuric acid.
34. Sodium sulfate + barium chloride (test for a sulfate).
35. Ionic equation for No. 34.
36. Sodium peroxide + water.
37. $\text{NaCl} + \text{H}_2\text{O} + \text{NH}_3 + \text{CO}_2$ (Solvay process).
38. Sodium bicarbonate (heated).
39. Hydrolysis reaction as in the reaction between:
 - (a) Sodium carbonate + water.
 - (b) Copper sulfate + water.
40. Preparation of pure nitrogen by heating ammonium nitrate.
41. Laboratory preparation of ammonia.
42. Ammonia + water.
43. Ammonia + hydrochloric acid.
44. Haber process for making ammonia.
45. Preparation of nitrous oxide by heating ammonium nitrate.
46. Laboratory preparation of nitric acid.
47. Copper + dilute nitric acid.
48. Copper + concentrated nitric acid.
49. Equations for arc process for making nitric acid.
50. Equations for manufacture of nitric acid from ammonia.
51. Equations for similar laboratory preparation of chlorine, bromine, and iodine.
52. Magnesium bromide + chlorine.
53. Equations for similar preparation of hydrofluoric acid, hydrochloric acid, and hydrobromic acid.
54. Reactions of carbon at temperature of electric furnace:
 - (a) Manufacture of calcium carbide.
 - (b) Manufacture of carborundum.
55. Calcium carbide + water.
56. Laboratory preparation of carbon dioxide.
57. Carbon dioxide + water.
58. Photosynthesis in green plants (carbon dioxide + water).

59. Carbon dioxide + limewater.
60. $\text{CaCO}_3 + \text{H}_2\text{O} + \text{CO}_2$ (formation of temporary hard water).
61. Baking powder reaction (sodium bicarbonate + cream of tartar).
62. Fire extinguisher reaction (sodium bicarbonate + sulfuric acid).
63. Preparation of carbon monoxide from carbon dioxide.
64. Burning of carbon monoxide.
65. Manufacture of water gas.
66. Softening of temporary hard water:
 - (a) Calcium bicarbonate (heated).
 - (b) Calcium bicarbonate + limewater.
67. Softening of permanent hard water ($\text{CaSO}_4 + \text{Na}_2\text{CO}_3$).
68. Manufacture of lime from limestone.
69. Slaking of lime.
70. Manufacture of bleaching powder.
71. Bleaching powder + hydrochloric acid.
72. Making plaster of Paris from gypsum.
73. Reduction of ferric oxide by carbon.
74. $\text{CaCO}_3 + \text{SiO}_2$ (formation of slag in the blast furnace).
75. Roasting of zinc sulfide ore.
76. Heating of zinc carbonate ore.
77. Reduction of zinc oxide by carbon.
78. Electrolysis of bauxite.
79. Thermit reaction ($\text{Fe}_2\text{O}_3 + \text{Al}$).
80. Aluminum + hydrochloric acid.
81. Aluminum + sodium hydroxide.
82. Preparation of aluminum hydroxide.
83. Burning of acetylene.
84. Burning of gasoline (assume formula is C_6H_{14}).
85. Conversion of starch to sugar.
86. Alcoholic fermentation of fruit sugar.
87. Acetic acid fermentation of alcohol.
88. Preparation of an ester ($\text{C}_2\text{H}_5\text{OH} + \text{HC}_2\text{H}_3\text{O}_2$).
89. Manufacture of soap [$\text{C}_3\text{H}_5(\text{C}_{18}\text{H}_{35}\text{O}_2)_3 + \text{NaOH}$].

90. Manufacture of nitroglycerine $[\text{C}_3\text{H}_5(\text{OH})_3 + \text{HNO}_3]$.
91. Inversion of sugar $(\text{C}_{12}\text{H}_{22}\text{O}_{11} + \text{H}_2\text{O})$.

Write electronic equations for the following reactions:

92. $\text{Na} + \text{Cl}_2$.
93. $\text{Zn} + \text{HCl}$.
94. $\text{Cu} + \text{AgNO}_3$.
95. $\text{NaOH} + \text{HCl}$.
96. $\text{KOH} + \text{H}_2\text{SO}_4$.
97. $\text{FeCl}_2 + \text{Cl}_2$.
98. $\text{FeCl}_3 + \text{H}_2$.
99. Electrolysis of brine.
100. Electrolysis of CuSO_4 solution.

UNIT 53. CHEMICAL ARITHMETIC

Atomic Weights

The foundation of chemical arithmetic is the table of the atomic weights of the elements begun by Dalton. Dalton realized that he could compare the weights of the atoms of other elements with the weight of an atom of one element chosen for a standard. He chose the lightest element, hydrogen, for his standard and called its atomic weight one. Since oxygen combines with more elements than hydrogen, it has since been found easier to use oxygen as a standard in determining the atomic weights of the elements. By making the relative atomic weight of oxygen 16, the weight of the lightest atom, that of hydrogen, remains approximately one (1.0078), and many of the other elements show atomic weights which are close to whole numbers.

A table of the atomic weights of the common elements will be found on the inside back cover. Even though it has now become possible to determine actual weights of the atoms, such figures, representing extremely minute quantities, are not convenient for use in most calculations. The actual weight of an atom of oxygen is about .000,000,000,000,000,000,000,026 gram; its relative weight is 16.

Type Problems

Examples of and practice material for the various type problems encountered in elementary chemistry follow:

TYPE 1. To find the molecular weight of a compound from its formula.

A chemical symbol represents more than the name of an element. It stands also for a definite amount of the element. It may represent one atom of the element or one atomic weight of the element. Thus, K represents one atom of potassium, or 39 grams (or any other weight unit) of potassium.

Similarly, a formula represents not only the name of a compound but also one molecule or one molecular weight of the

compound. (The molecular weight of a compound is the weight of one molecule of the compound as compared with the atomic weight of oxygen as 16.) To obtain the molecular weight of a compound we must add the weights of all the atoms in its molecule. Since the atomic weights used are only relative weights, it follows that the molecular weight obtained is also only a relative weight.

EXAMPLE A. Find the molecular weight of copper oxide, CuO.

Place the approximate atomic weights (refer to inside back cover) under the symbols and add them.

$$\begin{array}{ccc} \text{Cu} & \text{O} & \\ 63.6 & + 16 & = 79.6 = \text{mol. wt. of CuO.} \end{array}$$

EXAMPLE B. Find the molecular weight of calcium carbonate, CaCO₃.

Be sure to multiply the atomic weight of every element whose symbol is followed by a subscript by that subscript.

$$\begin{array}{ccc} \text{Ca} & \text{C} & \text{O}_3 \\ 40 + 12 + 16 \times 3 & = 40 + 12 + 48 & = 100 = \\ & \text{mol. wt. of CaCO}_3. & \end{array}$$

EXAMPLE C. Find the relative weight of 3 ZnSO₄ (zinc sulfate).

Be sure to multiply everything after a coefficient by the coefficient.

$$\begin{array}{l} 3 \text{ Zn} \quad \text{S} \quad \text{O}_4 \\ 3 (65.4 + 32 + 16 \times 4) \\ 3 (65.4 + 32 + 64) \\ 3 (161.4) = 484.2 = \text{relative weight of 3 molecules of ZnSO}_4. \end{array}$$

EXAMPLE D. Find the molecular weight of aluminum sulfate, Al₂(SO₄)₃.

Remember that the subscript multiplies only the element or radical whose symbol immediately precedes it.

$$\begin{array}{l} \text{Al}_2 \quad (\text{S} \quad \text{O}_4)_3 \\ 27 \times 2 + (32 + 16 \times 4)3 \end{array}$$

$$54 + (32 + 64)3$$

$$54 + (96)3$$

$$54 + 288 = 342 = \text{mol. wt. of } \text{Al}_2(\text{SO}_4)_3.$$

EXAMPLE E. Find the molecular weight of crystallized copper sulfate, $\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$.

Water of crystallization is chemically part of the compound and is separated from the rest of the formula by a dot only for the sake of convenience. The dot stands for $+$ and is not a multiplication sign.

$$\begin{array}{ccccccc} \text{Cu} & \text{S} & \text{O}_4 & \cdot & 5 & \text{H}_2 & \text{O} \\ 63.6 & + & 32 & + & 16 \times 4 & + & 5(1 \times 2 + 16) \end{array}$$

$$63.6 + 32 + 64 \quad + 5(18)$$

$$63.6 + 32 + 64 + 90 = 249.6 = \text{mol. wt. of crystallized copper sulfate.}$$

A very convenient unit in many calculations is the *gram-molecular weight* or *mole*. This is the molecular weight of the substance expressed in grams. Thus a mole of copper oxide, CuO (See Ex. A), is 79.6 grams.

Practice Work on Problems of Type 1

1. Calculate the molecular weight of sodium chloride.
2. Find the molecular weights of AgBr and CaO .
3. Calculate the molecular weights of MgSO_4 , Na_2CO_3 , and CaCl_2 .
4. What is the weight of a mole of sodium phosphate?
5. Find the molecular weights of $\text{Ca}(\text{HCO}_3)_2$, $\text{Al}(\text{NO}_3)_3$, and $\text{Ca}_3(\text{PO}_4)_2$.
6. Calculate the molecular weight of $\text{BaCl}_2 \cdot 2\text{H}_2\text{O}$.
7. What is the weight of a mole of gypsum, $\text{CaSO}_4 \cdot 2\text{H}_2\text{O}$?
8. Find the molecular weight of plaster of Paris, $(\text{CaSO}_4)_2 \cdot \text{H}_2\text{O}$.
9. The formula of Glauber's salt is $\text{Na}_2\text{SO}_4 \cdot 10\text{H}_2\text{O}$. Find the relative weight of 5 molecules of this substance.
10. Calculate the relative weight of $3\text{KAl}(\text{SO}_4)_2 \cdot 12\text{H}_2\text{O}$.

TYPE 2. To find the percentage composition of a compound from its formula.

To find the percentage composition of a compound is to find the per cent by weight of each different element present in the compound. To do this we divide the atomic weight of each element by the molecular weight of the compound and multiply the fraction thus obtained by 100.

EXAMPLE A. Find the percentage composition of potassium chlorate, $KClO_3$.

$$\begin{array}{r} K \quad Cl \quad O_3 \\ 39 + 35.5 + 16 \times 3 \end{array}$$

$$39 + 35.5 + 48 = 122.5 = \text{molecular weight of potassium chlorate.}$$

$$\% \text{ of potassium} = \frac{\text{atomic weight of K}}{\text{mol. wt. of } KClO_3} \times 100 = \frac{39 \times 100}{122.5} = 31.84\%.$$

$$\% \text{ of chlorine} = \frac{\text{atomic weight of Cl}}{\text{mol. wt. of } KClO_3} \times 100 = \frac{35.5 \times 100}{122.5} = 28.98\%.$$

$$\begin{array}{r} \% \text{ of oxygen} = \frac{\text{relative weight of } 3O}{\text{mol. wt. of } KClO_3} \times 100 = \frac{48 \times 100}{122.5} = 39.18\% \\ \hline \text{Total} \qquad \qquad \qquad = 100\% \end{array}$$

As a check on the problem, the total should be approximately 100.

EXAMPLE B. Find the per cent of water of crystallization in $CuSO_4 \cdot 5H_2O$.

$$\begin{array}{r} Cu \quad S \quad O_4 \quad \cdot 5H_2O \\ 63.6 + 32 + 16 \times 4 + 5 (1 \times 2 + 16) \end{array}$$

$$63.6 + 32 + 64 + 5 (18)$$

$$63.6 + 32 + 64 + 90 = 249.6 = \text{mol. wt. of } CuSO_4 \cdot 5H_2O.$$

$$\begin{array}{r} \% \text{ of water} = \frac{\text{rel. wt. of } 5H_2O}{\text{mol. wt. of crystal}} \times 100 \\ \qquad \qquad \qquad = \frac{5 (2 + 16)}{249.6} \times 100 = \frac{9000}{249.6} = 36.06\% \end{array}$$

EXAMPLE C. Find the percentage composition of $\text{Na}_2\text{SO}_4 \cdot 10\text{H}_2\text{O}$.

$$\begin{array}{r} \text{Na}_2 \quad \text{S} \quad \text{O}_4 \quad \cdot 10\text{H}_2\text{O} \\ 23 \times 2 + 32 + 16 \times 4 + 10(1 \times 2 + 16) \\ 46 + 32 + 64 + 180 = 322 = \text{mol. wt. of the compound.} \end{array}$$

Now find the total atomic weight of each element in the compound, thus:

$$\begin{array}{rcl} \text{sodium} & = & 2 \text{ atoms} = 23 \times 2 = 46 = \text{total atomic weight of Na} \\ \text{sulfur} & = & 1 \text{ atom} = 32 = \text{total atomic weight of S} \\ \text{oxygen} & = & 14 \text{ atoms} = 16 \times 14 = 224 = \text{total atomic weight of O} \\ \text{hydrogen} & = & 20 \text{ atoms} = 1 \times 20 = 20 = \text{total atomic weight of H} \\ \hline \text{Sum of the atomic weights} & = & 322 = \text{mol. wt. of the compound.} \\ \% \text{ of sodium} & = & \frac{46}{322} \times 100 = 14.3\% \\ \% \text{ of sulfur} & = & \frac{32}{322} \times 100 = 10.0\% \\ \% \text{ of oxygen} & = & \frac{224}{322} \times 100 = 69.5\% \\ \% \text{ of hydrogen} & = & \frac{20}{322} \times 100 = 6.2\% \\ \hline \text{Total} & = & 100\% \end{array}$$

EXAMPLE D. Find the weight of iron in 80 lbs. of an ore containing 90% ferric oxide, Fe_2O_3 .

$$\begin{array}{r} \text{Fe}_2 \quad \text{O}_3 \\ 56 \times 2 + 16 \times 3 = 160 = \text{mol. wt. of } \text{Fe}_2\text{O}_3. \\ \% \text{ of iron in } \text{Fe}_2\text{O}_3 = \frac{112}{160} \times 100 = 70\%. \end{array}$$

Now the weight of the iron oxide present in the ore is 90% of 80 lb. or 72 lb.

Therefore, 70% of 72 lb., or 50.4 lb., is the weight of the iron in the ore.

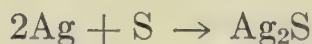
Practice Work on Problems of Type 2

1. Calculate the percentage composition of (a) water, (b) hydrogen chloride, and (c) zinc sulfate.
2. Calculate the percentage composition of (a) CaCO_3 , (b) HNO_3 , and (c) $(\text{CaSO}_4)_2 \cdot \text{H}_2\text{O}$.
3. Find the percentage composition of alum, $\text{KAl}(\text{SO}_4)_2 \cdot 12\text{H}_2\text{O}$.

4. Calculate the percentage composition of crystallized ferrous ammonium sulfate whose formula is $\text{FeSO}_4 \cdot (\text{NH}_4)_2\text{SO}_4 \cdot 6\text{H}_2\text{O}$.
5. Calculate the percent of oxygen in mercuric oxide (HgO).
6. Find the percent of water in crystals of barium chloride ($\text{BaCl}_2 \cdot 2\text{H}_2\text{O}$).
7. Calculate the amount of sulfur in 10 tons of sulfuric acid.
8. A zinc ore contains 80% of zinc oxide (ZnO). What weight of zinc is available from 100 tons of this ore?
9. Find the weight of aluminum oxide (Al_2O_3) that must be used in order to obtain 1 ton of aluminum.
10. Which ore will give more copper, one containing 95% of CuCO_3 or one containing 80% of Cu_2S ?
11. How much phosphorus can be obtained from 150 tons of bones containing 53% of $\text{Ca}_3(\text{PO}_4)_2$.

TYPE 3. Problems based on chemical formulas and equations.

Since the symbol of an element and the formula of a compound represent definite weights, an equation also represents definite weights of the substances taking part in the reaction. Thus



may be read, "2 atoms of silver plus 1 atom of sulfur yield 1 molecule of silver sulfide," or "216 grams of silver plus 32 grams of sulfur yield 248 grams of silver sulfide."

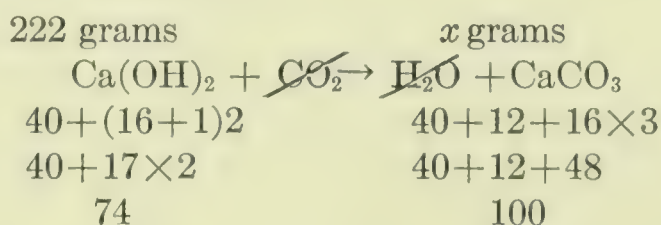
Problems based on chemical formulas and equations may be broadly divided into three groups:

- A. *Straight-weight problems* in which one weight is given and another is to be found.
- B. *Weight-volume problems* in which a weight is given and a volume is to be found, or *Volume-weight problems* in which a volume is given and a weight is to be found.
- C. *Straight-volume problems* in which one volume is given and another volume is to be found.

How to Solve Straight-Weight Problems (Type 3A)

1. Read the problem until you understand it thoroughly.
2. Write a balanced equation for the reaction involved. (The sum of the weights on both sides of the equation must, according to the law of the conservation of matter, be the same. We must bear in mind that even if we had no system of writing equations, the mathematical relationship of reactions would still hold).
3. Place the weight given *over* the formula of the substance involved. Place the weight required, designated by *x grams* (or any other unit of weight mentioned) *over* the formula of the substance whose weight is to be found.
4. Cross out all other formulas in the equation.
5. Since the same relationship exists between the actual weights expressed in grams, pounds, etc., as exists between the molecular weights represented by the equations, place the molecular weights of only those substances involved in the problem *under* the formulas of those substances. Do not ignore any coefficient.
6. Check the problem again and solve for *x*.

EXAMPLE. How many grams of calcium carbonate will be formed by the complete reaction between 222 grams of calcium hydroxide and carbon dioxide?



$$\frac{\text{weight of substance given}}{\text{mol. wt. of substance given}} = \frac{\text{weight of substance required}}{\text{mol. wt. of substance required}}$$

$$\frac{222}{74} = \frac{x}{100}$$

$$\begin{aligned} \text{Solving for } x, \quad 22,200 &= 74x \\ x &= 300 \end{aligned}$$

Therefore, 300 grams of calcium carbonate will be formed.

Alternate method. We can avoid the use of an equation involving x by solving the problem as follows:

$$\frac{\text{wt. of substance given}}{\text{mol. wt. of subst. given}} \times \text{mol. wt. of subst. required} = \text{Ans.,}$$

$$\frac{222}{74} \times 100 = 300.$$

Practice Work on Problems of Type 3A

1. How much zinc is required to react with sufficient sulfuric acid to produce 8 grams of hydrogen?

2. How much oxygen can be obtained from 100 grams of KClO_3 ?

3. How much HCl reacting with MnO_2 would be needed to prepare 100 grams of chlorine.

4. What weight of carbon dioxide can be obtained by burning 10 tons of carbon?

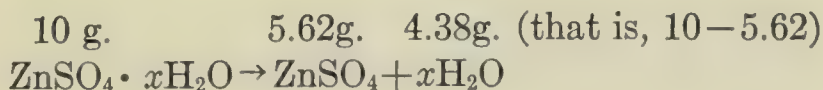
5. Ten grams of sodium completely react with water. What weight of sodium hydroxide is formed? What weight of hydrogen is liberated?

6. Twenty ounces of hydrogen were liberated by the electrolysis of water. What weight of oxygen was formed at the same time?

7. What weight of magnesium will be needed to react with hydrochloric acid to produce 80 grams of MgCl_2 ? What weight of hydrogen will be evolved?

8. Ten grams of crystallized zinc sulfate gave, after heating, 5.62 grams of the anhydrous salt, ZnSO_4 . Calculate the number of molecules of water of crystallization in the original compound.

Let x represent the number of molecules of water of crystallization. Then



Now complete the problem.

9. Seven grams of crystalline calcium sulfate gave 5.536 grams of the anhydrous salt after being heated. Calculate the number of molecules of water of crystallization.

10. Ten grams of crystalline magnesium sulfate were made from 4.88 grams of anhydrous MgSO_4 . What is the formula of the crystalline salt?

11. If a sample of coal contains 80% of carbon, how many pounds of carbon dioxide will be produced when one ton (2000 lbs.) of this coal is burned? What weight of oxygen is required for this combustion? Assuming air to contain 23% by weight of oxygen, how many pounds of air are used up?

12. If a student placed 8 grams of zinc, and dilute acid containing 9.8 grams of sulfuric acid in his hydrogen generator, which of these two substances would be in excess? How much of it?

13. How much iron can be obtained from 150 tons of hematite ore which contains 85% of ferric oxide?

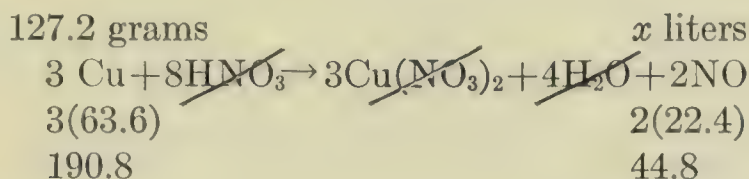
TYPE 3B. Weight-volume and volume-weight problems.

A study of a large number of gases and vapors has led to the conclusion that the molecular weight of a gas or vapor expressed in grams always occupies 22.4 liters under standard conditions of temperature and pressure. For example, under standard conditions 2 grams of hydrogen (mol. wt. = 2) occupies 22.4 liters, 32 grams of oxygen (mol. wt. = 32) occupies 22.4 liters, 17 grams of ammonia (mol. wt. = 17) occupies 22.4 liters, 44 grams of carbon dioxide (mol. wt. = 44) occupies 22.4 liters. This volume of 22.4 liters is known as the *gram-molecular-volume*.

In solving weight-volume and volume-weight problems, the same relationship holds as in straight-weight problems, except that we may now substitute for the gram-molecular-weight of the substance whose volume is to be found its gram-molecular-volume of 22.4 liters for each molecule of the substance.

In all problems of this type it is assumed that the reaction is taking place under standard conditions of temperature and pressure.

EXAMPLE A. How many liters of nitric oxide can be prepared by the action of sufficient dilute nitric acid on 127.2 grams of copper?



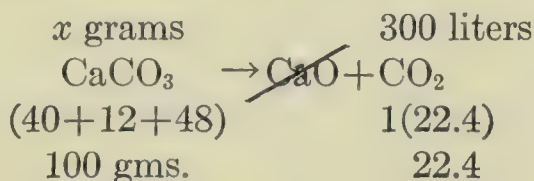
$$\frac{\text{weight of subst. given}}{\text{mol. wt. of subst. given}} = \frac{\text{volume of subst. required}}{\text{gram-molecular-volume of subst. required}}$$

$$\frac{127.2}{190.8} = \frac{x \text{ liters}}{44.8}$$

$$190.8x = 127.2(44.8)$$

$$x = 29.9 \text{ liters}$$

EXAMPLE B. How many grams of calcium carbonate must be heated to produce 300 liters of carbon dioxide gas?



$$\frac{x}{100} = \frac{300}{22.4}$$

$$22.4x = 100 \times 300$$

$$x = \frac{100 \times 300}{22.4}$$

$$x = 1339.29 \text{ grams. of CaCO}_3$$

EXAMPLE C. What is the volume under standard conditions of 50 grams of ammonia gas?

$$\begin{array}{l}
 \text{NH}_3 \\
 \text{mol. wt.} = 14 + 3 = 17 \\
 50 \text{ g.} = x \text{ liters} \\
 (\text{mol. wt.}) 17 \text{ g.} = 22.4 \text{ liters}
 \end{array}$$

$$\frac{50}{17} = \frac{x}{22.4}$$

$$17x = 50 \times 22.4$$

$$x = \frac{50 \times 22.4}{17}$$

$$= 65.88 \text{ liters of ammonia.}$$

EXAMPLE D. Find the weight of 500 liters of carbon dioxide under standard conditions.



$$\text{mol. wt.} = 12 + 32 = 44$$

$$x \text{ g.} = 500 \text{ liters}$$

$$(\text{mol. wt.}) 44 \text{ g.} = 22.4 \text{ liters}$$

$$\frac{x}{44} = \frac{500}{22.4}$$

$$22.4x = 44 \times 500$$

$$x = \frac{44 \times 500}{22.4} = 982.14 \text{ grams of CO}_2$$

Practice Work on Problems of Type 3B

1. A piece of pure carbon weighing 30 grams was burned. What volume of carbon dioxide was formed?
2. How many liters of hydrogen gas can be formed by the action of 36 grams of sodium on sufficient water?
3. How much marble (CaCO_3) will be needed to produce 500 liters of carbon dioxide gas, assuming sufficient hydrochloric acid is used?
4. What weight of copper must be dissolved by concentrated sulfuric acid to produce 58.5 cc. of sulfur dioxide gas?
5. A manufacturer requires 10,000 liters of nitrous oxide, N_2O . What weight of ammonium nitrate will have to be decomposed?
6. What volume of hydrogen sulfide gas can be obtained when 25 grams of ferrous sulfide reacts with sufficient hydrochloric acid?

7. Assuming air is one-fifth oxygen by volume, how many liters of air are needed to burn 500 grams of sulfur?

8. How many liters of carbon dioxide are produced by the complete combustion of 1500 grams of coal containing 90% of carbon?

9. How many liters of hydrogen chloride can be obtained from 5 grams of sodium chloride reacting with sufficient sulfuric acid?

10. Find the weight of 1 liter of hydrogen gas at standard conditions.

11. What is the weight of 100 liters of a gas that has a molecular weight of 16?

12. What is the volume, at standard conditions, of 1 kilogram (1000 grams) of hydrogen chloride gas?

13. Find the weight of 100 cc. (0.1 liter) of oxygen at standard conditions.

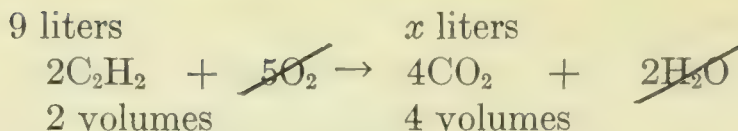
14. What is the molecular weight of a gas, 52 grams of which occupy 44.8 liters at standard conditions?

15. 200 cc. of a gas weigh 0.304 grams. Find its molecular weight.

TYPE 3C. Straight-volume problems.

Since the same volumes of gases under the same conditions of temperature and pressure contain the same number of molecules (Avogadro's law), we may conclude that the coefficients of formulas representing gases also represent their volumes. Therefore, the same relationships exist between the gram-molecular-weights of gases, their gram-molecular-volumes, and their coefficients. In the solution of problems involving only the volumes of gases (or volatile substances), we need consider only their volumes as expressed by the coefficients.

EXAMPLE A. How many liters of carbon dioxide will be formed during the complete combustion of 9 liters of acetylene gas (C_2H_2)?



$$\frac{\text{volume of subst. given}}{\text{coefficient of subst. given}} = \frac{\text{volume of subst. required}}{\text{coefficient of subst. required}}$$

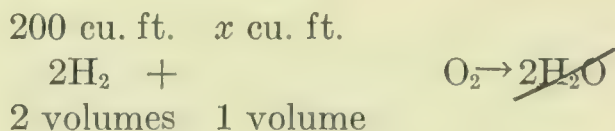
$$\frac{9}{2} = \frac{x}{4}$$

$$2x = 36$$

$$x = 18 \text{ liters of CO}_2$$

EXAMPLE B. What volume of air is needed for the complete combustion of 200 cu. ft. of hydrogen? Assume that air is one-fifth oxygen by volume.

(We first find the volume of oxygen and then compute the volume of air)



$$\frac{200}{2} = \frac{x}{1}$$

$$2x = 200$$

$$x = 100 \text{ cu. ft. of oxygen}$$

$$5 \times 100 = 500 \text{ cu. ft. of air}$$

Practice Work on Problems of Type 3C

1. Fifty liters of hydrogen react completely with nitrogen. What volume of ammonia gas is formed?
2. What volume of oxygen will convert sufficient sulfur into 150 cc. of sulfur dioxide?
3. What volume of hydrogen chloride gas can be made from 5000 cubic feet of chlorine?
4. How many liters of oxygen will be used during the complete combustion of 500 cc. of methane, CH₄? What volume of carbon dioxide will be produced?

5. Assume gasoline to have the formula C_6H_{14} . What volume of air is needed to completely burn 100 liters of gasoline vapor?

6. When some hydrogen sulfide gas was completely burned in air, 1000 liters of air were used up. What volume of hydrogen sulfide was burned?

7. What volume of air will convert 250 cc. of nitric oxide, NO, into nitrogen peroxide, NO_2 ?

TYPE 4. To determine the simplest formula of a compound from its percentage composition.

If we know by analysis the percentage composition of a compound, it is possible to determine the number of atoms of each element in a molecule of the compound and thus write a formula for the compound.

EXAMPLE. Find the simplest formula of a compound which contains 63.6% nitrogen and 36.4% oxygen.

$$\left. \begin{array}{l} \text{For N, } \frac{\% \text{ of N in compound}}{\text{atomic weight of N}} = \frac{63.6}{14} = 4.54 \\ \text{For O, } \frac{\% \text{ of O in compound}}{\text{atomic weight of O}} = \frac{36.4}{16} = 2.28 \end{array} \right\} \begin{array}{l} \text{Highest} \quad \frac{4.54}{2.28} = 2 \\ \text{Common} \\ \text{Factor} \quad \frac{2.28}{2.28} = 1 \\ \text{(H. C. F.)} \\ \quad = 2.28 \end{array}$$

The numbers 2 and 1 represent the smallest number of atoms of nitrogen and oxygen that can be present in a molecule of the compound. The simplest formula is, therefore, N_2O ; the true formula may be written $(N_2O)_x$.

Problems of this type can be checked by finding the percentage composition of the formula as found. This percentage composition should, of course, agree with the data given.

Practice Work on Problems of Type 4

1. A certain substance has the following percentage composition: barium, 58.8%; sulfur, 13.75%; and oxygen, 27.45%. Find its simplest formula.

2. Calculate the formula of the compound with the follow-

ing composition: calcium, 29.4%; sulfur, 23.56%, oxygen, 47.04%.

3. Find the simplest formula of a compound that contains: silver, 70.13%; nitrogen, 9.09%; oxygen, 20.77%.

4. What is the simplest formula of a compound which on analysis showed the following composition: carbon, 54.55%; hydrogen, 9.09%; and oxygen, 36.36%?

5. There are 1.6 grams of oxygen in 8.47 grams of an oxide of barium. Calculate the formula of this oxide.

6. Calculate the formula of magnesium chloride if 10 grams of it contains 7.443 grams of chlorine.

7. Bromine was allowed to act upon hot copper until it changed 32 grams of the metal into 112.5 grams of copper bromide. What is the simplest formula of this compound?

TYPE 5. To find the *true* formula of a compound.

In the case of compounds that are not solids, the simplest formula is not always the true formula. Very often the true formula is a simple multiple of the simple formula. If the simple formula is N_2O the true formula may be written $(N_2O)_x$. To find the value of x , we must know the molecular weight of the compound, for, by comparison of the molecular weight based on the simplest formula and the true molecular weight, we can find the value of x . For example, if the true molecular weight is twice the molecular weight based on the simplest formula, we know that there must be twice as many atoms of each element as the simplest formula gives or, in other words, that the value of x is 2.

We can find the molecular weight of a gas if we know either the *vapor density* of the gas or the weight of one liter of the gas, or it can be ascertained by certain experimental methods. The vapor density (V. D.) of a gas is the number of times that gas is as heavy as an equal volume of hydrogen, both gases being weighed under standard conditions. It has been experimentally determined that the vapor density of a gas is always equal to half its molecular weight; that is: $\text{Molecular weight} = 2\text{V.D.}$

Since the weight of one liter of hydrogen is 0.09 grams, the following formula also holds:

$$V. D. = \frac{\text{Weight in grams of 1 liter of the gas}}{0.09}$$

EXAMPLE A. Find the true formula of a compound which contains 63.6% nitrogen and 36.4% oxygen, and whose vapor density is 21.9%.

We find that, when solved in accordance with the method of Type 4, the simplest formula of this compound is N_2O . The molecular weight based on this simplest formula is $14 \times 2 + 16$, or 44.

Using the vapor density of the compound, we find that the true molecular weight is: $\text{Mol. wt.} = 2V.D. = 2(21.9) = 43.8$.

Since the true molecular weight (43.8) is approximately the same as the molecular weight based on the simplest formula, we know that the simplest formula is the true formula.

Had the true molecular weight been a multiple of the molecular weight based on the simplest formula, that is, x times it, we should have multiplied the simplest formula by this number x to get the true formula.

EXAMPLE B. What is the true formula of a compound which contains 92.3% carbon and 7.7% hydrogen, and 2.2 grams of its vapor occupy 628 cc. at standard conditions?

$$\begin{array}{l} \text{For C, } \frac{\% \text{ of C}}{\text{at. wt. of C}} = \frac{92.3}{12} = 7.69 \\ \text{For H, } \frac{\% \text{ of H}}{\text{at. wt. of H}} = \frac{7.7}{1} = 7.7 \end{array} \quad \left\{ \begin{array}{l} \frac{7.69}{7.69} = 1 = \text{smallest no. of C} \\ \text{atoms in the molecule.} \\ \text{H. C. F.} = 7.69 \\ \frac{7.7}{7.69} = 1 = \text{smallest no. of H} \\ \text{atoms in the molecule.} \end{array} \right.$$

The simplest formula, therefore, is CH . The true formula may be represented $(CH)_x$. The molecular weight based on this formula is $(12+1)x$, or $13x$.

Since 628 cc. weigh 2.2 grams, one liter (1000 cc.) will weigh:

$$\frac{1000}{628} \times 2.2 = 3.5 \text{ grams} = \text{weight of 1 liter}$$

Substituting this value in the equation, we obtain:

$$\text{V. D.} = \frac{3.5}{0.09} = 38.9 = \text{vapor density of the compound.}$$

Hence the true molecular weight $= 2 \text{ V. D.} = 2(38.9) = 77.8$

But this true molecular weight is $\frac{77.8}{13}$ or approximately 6 times as great as the molecular weight based on the simplest formula. Therefore the true formula is $(\text{CH})_6$, or C_6H_6 (benzene).

Practice Work on Problems of Type 5

1. A compound has the composition: carbon, 75%; hydrogen, 25%. Its molecular weight is 16. Find its formula.

2. A substance composed of carbon, 39.95%; hydrogen, 6.69%; and oxygen, 53.36%, has a molecular weight of 60. Calculate its formula.

3. The percentage composition of a certain gas is: nitrogen, 46.6%; oxygen, 53.4%. 100 cc. of this gas weigh 0.134 grams at standard conditions. Find its formula.

4. The vapor density of a substance is 21.9. It contains 27.4% carbon and 72.6% oxygen. What is its true formula?

5. A compound contains 24.24% carbon, 4.04% hydrogen, and 71.72% chlorine. Its vapor density is 49.6. Find its true formula.

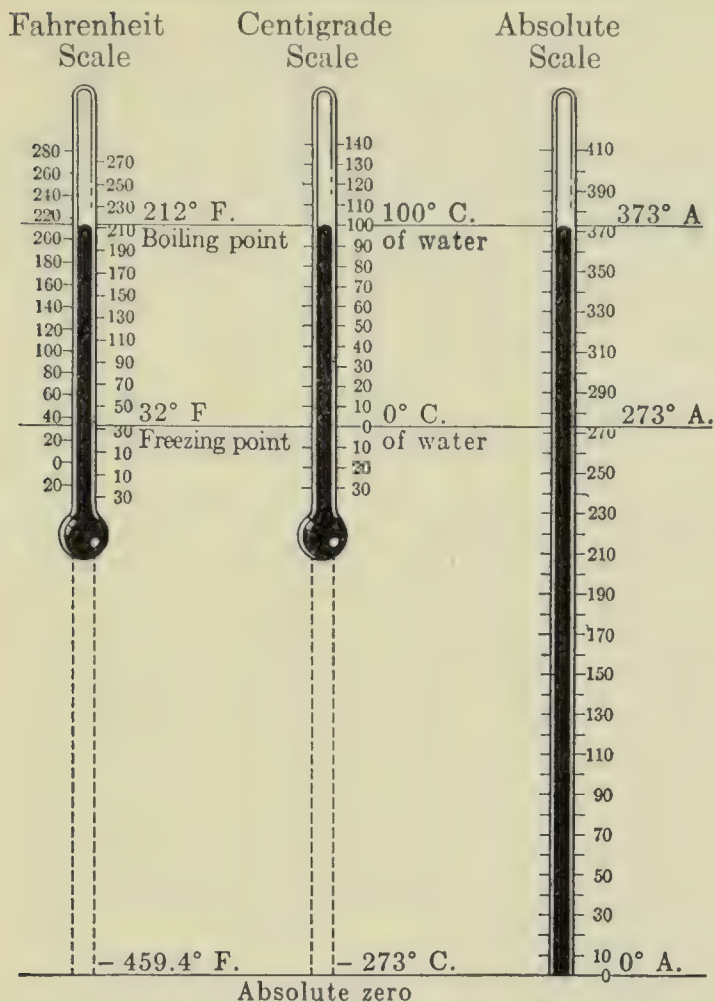
6. A compound has the following composition: carbon, 37.6%; hydrogen, 12.5%; and oxygen, 49.9%. The weight of 1 liter of this vapor is 1.434 grams. What is the true formula?

7. 57 cc. of a certain gas weighed 0.134 grams. Its composition was found to be 46.15% carbon and 53.85% nitrogen. Find its true formula.

TYPE 6. To convert thermometer readings from one scale to another.

1. To convert Centigrade readings to the corresponding Fahrenheit readings, multiply the Centigrade reading by 1.8 and add 32 degrees.

COMPARATIVE THERMOMETER SCALES



EXAMPLE. What is the Fahrenheit reading corresponding to $20^{\circ}\text{C}.$?

$$20 \times 1.8 + 32 = 36 + 32 = 68^{\circ}\text{F.}$$

2. To convert Fahrenheit readings to the corresponding Centigrade readings, subtract 32 degrees from the Fahrenheit reading and divide by 1.8.

EXAMPLE. What is the Centigrade reading corresponding to $100^{\circ}\text{F}.$?

$$\frac{100 - 32}{1.8} = \frac{68}{1.8} = 37.77^{\circ}\text{C.}$$

3. To convert Centigrade readings to the corresponding Absolute scale readings, add 273° to the Centigrade readings.

EXAMPLE. What is the Absolute scale reading corresponding to $-43^{\circ}\text{C}.$?

$$-43^{\circ} + 273^{\circ} = 230^{\circ}\text{ Absolute}$$

4. To convert Fahrenheit readings to Absolute scale readings, first convert the Fahrenheit reading to the Centigrade reading, and then add 273° .

EXAMPLE. What is the Absolute scale reading corresponding to $68^{\circ}\text{F}.$?

$$\frac{68 - 32}{1.8} + 273 = 20 + 273 = 293^{\circ}\text{ Absolute}$$

Practice Work on Problems of Type 6

1. Normal room temperature is $68^{\circ}\text{F}.$ To what temperature does this correspond on the Centigrade scale?

2. The temperature of liquid air is $-192^{\circ}\text{C}.$ To what temperature does this correspond on the Fahrenheit scale? On the Absolute scale?

3. The oxyacetylene flame is at a temperature of $3500^{\circ}\text{C}.$ To what temperature Absolute does this correspond? To what temperature Fahrenheit?

4. The melting point of aluminum is $1220^{\circ}\text{F}.$ To what temperature Centigrade does this correspond? To what temperature Absolute?

5. The freezing point of alcohol is $-202^{\circ}\text{F}.$ To what temperature Centigrade does this correspond? To what temperature Absolute?

TYPE 7. Problems involving Boyle's Law and Charles' Law.

Boyle's law and Charles' law may be expressed as the following equations, in which p represents the initial pressure, p' the final pressure, v the initial volume, v' the final volume, T the initial absolute temperature, and T' the final absolute temperature.

$$\text{Boyle's Law: } pv = p'v'$$

$$\text{Charles' Law: } \frac{v}{T} = \frac{v'}{T'}$$

By using these equations, gas measurements can be changed to suit the conditions wanted. Usually, they are used to change gas measurements from existing conditions to standard conditions.

A. Change in volume due to change in pressure. (Boyle's Law)

Example. A quantity of chlorine gas occupies a volume of 50 cc. when the barometer reads 740 mm. What would its volume be under standard conditions?

The temperature is not mentioned, so we assume it does not change.

$$\begin{aligned} pv &= p'v' \\ 740(50) &= 760(x) \\ 37,000 &= 760x \\ x &= 48.7 \text{ cc. of chlorine} \end{aligned}$$

B. Change in volume due to change in temperature. (Charles' Law)

Example. A quantity of carbon dioxide gas measured 450 cc. when the room temperature was 22° C. What would its volume be under standard conditions?

The pressure is not mentioned, so we assume we are working under standard or 760 mm. pressure.

$$\begin{aligned} \frac{v}{T} &= \frac{v'}{T'} \\ \frac{450}{273 + 22} &= \frac{x}{273} \\ \frac{450}{295} &= \frac{x}{273} \\ 295x &= 450(273) \\ x &= 416.4 \text{ cc. of CO}_2 \end{aligned}$$

C. Change in volume due to simultaneous change in temperature and pressure.

Example. At 27° C. and 800 mm. pressure a quantity of nitrogen gas measured 20 cc. What would the volume be under standard conditions?

$$\frac{pv}{T} = \frac{p'v'}{T'}$$

$$\frac{800(20)}{273+27} = \frac{760(x)}{273}$$

$$800(20)(273) = 760(300)x$$

$$x = \frac{800(20)(273)}{760(300)}$$

$$x = 19.2 \text{ cc. nitrogen}$$

Practice Work on Problems of Type 7

1. The volume of a sample of air is 20 liters at 20° C. What would its volume be at 60° C., the pressure remaining the same?

2. The volume of a sample of hydrogen is 300 cc. at 760 mm. pressure. What would its volume be at 740 mm. pressure, the temperature remaining the same?

3. The air within a half-inflated balloon occupies a volume of 200,000 liters at 10° C. temperature and 750 mm. pressure. What will be its volume after ascending to a point where the pressure is 350 mm. and the temperature is -10° C.?

4. What is the volume at standard conditions of 100 cc. of a gas which is at 21° C. temperature and 740 mm. pressure?

5. The air in an automobile tire is at a temperature of 20° C. and at a pressure of 3 kg. per square centimeter. Assuming the volume of the tire (and air) does not change, to what temperature has this air risen when it exerts a pressure of 4 kg. per square centimeter?

UNIT 54. BIOGRAPHY

Indicate the nationality and the contributions made to chemistry by the following distinguished chemists.

- | | |
|---------------------------------|-----------------------------|
| 1. Acheson (1856-) | 22. Gay-Lussac (1778-1850) |
| 2. Anderson, Carl D. (1905-) | 23. Haber (1868-1934) |
| 3. Arrhenius (1859-1927) | 24. Hall (1863-1914) |
| 4. Aston (1877-) | 25. Langmuir (1881-) |
| 5. Avogadro (1776-1856) | 26. Lavoisier (1743-1794) |
| 6. Baekeland (1863-) | 27. Lewis (1875-) |
| 7. Becquerel, Henri (1852-1908) | 28. Mendeleyeff (1834-1907) |
| 8. Berzelius (1779-1848) | 29. Millikan (1868-) |
| 9. Bohr (1885-) | 30. Moissan (1852-1907) |
| 10. Boyle (1627-1691) | 31. Moseley (1887-1915) |
| 11. Bunsen (1811-1899) | 32. Ostwald (1853-1932) |
| 12. Cavendish (1731-1810) | 33. Priestley (1733-1804) |
| 13. Chadwick, James (1891-) | 34. Ramsay (1852-1916) |
| 14. Charles (1746-1823) | 35. Rutherford, E. (1871-) |
| 15. Crookes (1832-1919) | 36. Scheele (1742-1786) |
| 16. Curie, Marie (1867-1934) | 37. Soddy (1877-) |
| 17. Dalton (1766-1844) | 38. Thomson, J. J. (1856-) |
| 18. Davy (1778-1829) | 39. Urey (1893-) |
| 19. De Broglie (1892-) | 40. Wilson, C.T.R. (1869-) |
| 20. Faraday (1791-1867) | 41. Woehler (1800-1882) |
| 21. Frasch (1852-1914) | |

UNIT 55. TABLE OF COMMON SUBSTANCES AND ALLOYS

I. In tabular form indicate the chemical name, formula and uses (2 if possible) for each of the following:

- | | |
|-------------------------|----------------------|
| 1. Alcohol (grain) | 30. Gypsum |
| 2. Alcohol (wood) | 31. Hematite |
| 3. Alum | 32. Hypo |
| 4. Ammonia water | 33. Iron pyrites |
| 5. Aqua regia | 34. Laughing gas |
| 6. Baking powder | 35. Lime |
| 7. Baking soda | 36. Limewater |
| 8. Bleaching powder | 37. Lithopone |
| 9. Blue vitriol | 38. Lunar caustic |
| 10. Borax | 39. Marble |
| 11. Calomel | 40. Marsh gas |
| 12. Carbolic acid | 41. Milk of magnesia |
| 13. Carborundum | 42. Muriatic acid |
| 14. Carnallite | 43. Oil of vitriol |
| 15. Caustic potash | 44. Plaster of Paris |
| 16. Caustic soda | 45. Quartz |
| 17. Chile saltpeter | 46. Rochelle salt |
| 18. Cinnabar | 47. Rock phosphate |
| 19. Corrosive sublimate | 48. Salt (common) |
| 20. Corundum | 49. Sal ammoniac |
| 21. Cream of tartar | 50. Saltpeter |
| 22. Dakin's solution | 51. Slaked lime |
| 23. Epsom salts | 52. Soap (hard) |
| 24. Ether | 53. Soap (soft) |
| 25. Fluorspar | 54. Vinegar |
| 26. Galena | 55. Washing soda |
| 27. Glass (common) | 56. Water glass |
| 28. Glucose | 57. White lead |
| 29. Glycerine | 58. Zinc white |

II. In tabular form indicate the composition and uses (2 if possible) for each of the following alloys:

- | | |
|---------------------|---------------------|
| 1. Aluminum bronze | 15. Nickel coin |
| 2. Babbitt metal | 16. Nickel steel |
| 3. Bell metal | 17. Nichrome |
| 4. Brass | 18. Pewter |
| 5. Bronze | 19. Silicon steel |
| 6. Copper coin | 20. Silver amalgam |
| 7. Duralumin | 21. Silver coin |
| 8. Duriron | 22. Solder |
| 9. German silver | 23. Stainless steel |
| 10. Gold coin | 24. Tungsten steel |
| 11. Invar steel | 25. Type metal |
| 12. Magnalium | 26. Vanadium steel |
| 13. Manganese steel | 27. Wood's metal |
| 14. Monel metal | |

UNIT 56. REVIEW OF IMPORTANT TERMS AND LAWS

I. In tabular form indicate for each of the following terms its definition (or explanation), and an example (or illustration):

- | | |
|------------------------------|-------------------------------------|
| 1. Acid | 32. Electrolyte |
| 2. Acid anhydride | 33. Electron |
| 3. Acid salt | 34. Element |
| 4. Alcohol | 35. Emulsion |
| 5. Alkali | 36. Ester |
| 6. Allotropism | 37. Fermentation |
| 7. Alloy | 38. Filtration |
| 8. Amalgam | 39. Fixation of nitrogen |
| 9. Amorphous | 40. Flux |
| 10. Amphoteric element | 41. Fractional distillation |
| 11. Analysis | 42. Gram molecular volume |
| 12. Atom | 43. Hard water |
| 13. Atomic weight | 44. Heat of formation of a compound |
| 14. Base | 45. Hydrate |
| 15. Base anhydride | 46. Hydrocarbon |
| 16. Basic salt | 47. Hydrolysis |
| 17. Carbohydrate | 48. Hygroscopic substance |
| 18. Catalyst | 49. Inversion of sugar |
| 19. Chemical change | 50. Ion |
| 20. Colloid | 51. Isomers |
| 21. Combustion | 52. Isotopes |
| 22. Compound | 53. Kindling temperature |
| 23. Decomposition | 54. Malleability |
| 24. Deliquescent substance | 55. Mineral |
| 25. Destructive distillation | 56. Miscible liquids |
| 26. Distillation | 57. Molar solution |
| 27. Double salt | 58. Molecular weight |
| 28. Ductility | 59. Molecule |
| 29. Effervescence | 60. Mordant |
| 30. Efflorescent substance | 61. Neutralization |
| 31. Electrolysis | |

- | | |
|-------------------------|------------------------------|
| 62. Neutron | 77. Saturated solution |
| 63. Normal salt | 78. Slag |
| 64. Normal solution | 79. Solution |
| 65. Occlusion | 80. Spontaneous combustion |
| 66. Ore | 81. Sublimation |
| 67. Oxidation | 82. Supersaturated solution |
| 68. Oxidizing agent | 83. Suspension |
| 69. Positron | 84. Synthesis |
| 70. Precipitate | 85. Tempering |
| 71. Proton | 86. Tincture |
| 72. Reducing agent | 87. Valence |
| 73. Relative humidity | 88. Vapor density |
| 74. Reversible reaction | 89. Water of crystallization |
| 75. Salt | 90. Welding |
| 76. Saponification | |

II. In tabular form indicate for each of the following laws, a brief statement of the law, and an illustration (or example).

- | | |
|---------------------------|--------------------------------|
| 1. Avogadro's Law | 6. Law of Conservation of |
| 2. Boyle's Law | Matter |
| 3. Charles' Law | 7. Law of Definition Propor- |
| 4. Gay-Lussac's Law | tions |
| 5. Law of Conservation of | 8. Law of Mass Action |
| Energy | 9. Law of Multiple Proportions |

Optional Questions

1. How are the elements arranged according to the periodic table of Mendeleeff?
2. State the periodic law. What is meant by a period? A family of elements?
3. Explain the difference between Moseley's arrangement of the elements, and Mendeleeff's arrangement.
4. What uses have such arrangements of the elements?
5. Explain the law of mass action.
6. What is Prout's hypothesis?
7. What were the assumptions of Dalton's atomic theory? In what respects has the electron theory changed Dalton's theory?

Concentrations of Solutions Used in the Experiments

1. The solutions called for in the experiments are 10% solutions unless otherwise indicated either in the experiment or in the list of apparatus and materials for the experiments which begins on page 228.

2. In many cases, solutions of 5% or less strength will work as well as 10% solutions. Experimentation by the teacher will determine most satisfactorily the deviations from 10% strength which are usable in his laboratory.

3. For the average class in chemistry, it is wise to make up stock solutions of the more commonly used substances in 2½ liter acid bottles.

4. Silver nitrate solution should be about 1% strength.

5. Solutions of acids.

(a) Dilute hydrochloric acid: Add 1 part of concentrated acid to 4 parts of water.

(b) Dilute sulfuric acid: Add 1 part of concentrated acid to 5 parts of water in an earthenware vessel.

(c) Dilute nitric acid: Add 1 part of concentrated acid to 4 parts of water.

(d) Dilute acetic acid: Add 1 part of glacial acetic acid to 10 parts of water.

6. Solutions of bases.

(a) Dilute sodium hydroxide: 10% solution.

(b) Dilute potassium hydroxide: 10% solution.

(c) Dilute ammonium hydroxide: Add 1 part of concentrated ammonium hydroxide to 4 parts of water.

(d) Concentrated sodium hydroxide: 30% solution.

(e) Concentrated potassium hydroxide: 30% solution.

(f) Limewater (calcium hydroxide). Put about 3 tablespoonfuls of powdered calcium oxide into a 2½ liter acid bottle. Fill with water and shake thoroughly. Allow to settle. The clear liquid when decanted or siphoned off is a saturated solution of

calcium hydroxide. By refilling the bottle with water, and shaking again, limewater can thus be made many times.

7. Fehling's Solution (Two separate solutions):

A. 34.66 grams of $\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$ in 500 cc. of solution

B. 173 grams of $\text{KNaC}_4\text{H}_4\text{O}_6 \cdot 4\text{H}_2\text{O}$ }
50 grams of NaOH } in 500 cc. of solution

For use, mix equal volumes of solutions A and B at time of using.

8. Benedict's Solution (One solution):

(1) Sodium citrate—200 grams	} Dissolve in	
Potassium thiocyanate—125 grams		} 700 cc. of
Anhydrous sodium carbonate—100 grams		

(2) Dissolve 18 grams of $\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$ in 100 cc. of water.

NOTE: *Unless otherwise indicated in the experiment, dilute acids and bases are to be used.*

NOTE: *In making up stock solutions, it should be remembered that 1 cc. of water weighs approximately 1 gram, at room temperature.*

Apparatus and Materials

For convenience, the apparatus and materials required in each unit are listed below. Notice that the additional apparatus and materials required for demonstrations are listed separately.

Introductory Unit: How We Learn, Think, and Solve Problems.

Apparatus. Glass jet tube, metal pan.

Materials. Candle, wire gauze, stick of wood, small stone, water, twigs or leaves, matches.

Unit 1. Physical and Chemical Changes in Matter

Apparatus. Bunsen burner, evaporating dish, forceps, glass rod, 3 test tubes, watch glass, test tube holder.

Materials. Copper wire, platinum wire (one end sealed in a glass rod), magnesium ribbon, granulated sugar, concentrated sulfuric acid.

Unit 2. Elements, Compounds, and Mixtures

Apparatus. Bunsen burner, magnet, test tube holder, 4 test tubes, beaker, spatula.

Materials. Powdered sulfur, iron filings, carbon disulfide, wood splints.

Unit 3. The Explanation of Burning (See diagram, page 7.)

Apparatus. Ring stand and ring, Bunsen burner, horn pan balance with metal pan, asbestos paper, test tube holder, hard glass test tube. (*Alternate procedure.* Crucible and cover, pipestem triangle, tripod, forceps, beam balance.)

Materials. Powdered (or ribbon) magnesium, sand, iron filings, wood splints, mercuric oxide, mercury in glass container for observation.

Unit 4. Oxygen (See diagram, page 11.)

Apparatus. Ring stand, clamps, Bunsen burner, test tube fitted with a one-hole rubber stopper, delivery tube, water trough, 4 collecting bottles, 4 glass plates, forceps, combustion spoon.

Materials. Potassium chlorate, manganese dioxide, wood splints, limewater, pieces of charcoal, sulfur, and red phosphorus.

Unit 5. Hydrogen (See diagram, page 15.)

Apparatus. Generating bottle fitted with a two-hole rubber stopper, thistle tube, delivery tube, water trough, 4 collecting bottles, 4 glass plates, Bunsen burner.

Materials. Mossy zinc, dilute sulfuric acid, wood splints, candle.

Demonstration. 2 ring stands, 1 ring, 2 clamps, wire gauze, 2 Bunsen burners, fish-tail burner, filter paper, funnel, beaker, watch glass, wire gauze, hydrogen generator, calcium chloride, drying tube, clamp, fish-tail burner, hard glass tube; copper oxide (wire form preferable). (See diagram, page 16.)

Unit 6. Action of Metals with Water

Apparatus. Bunsen burner, test tube rack, test tube holder, 4 test tubes, water trough, one-hole rubber stopper to fit test tube, delivery tube.

Materials. Wood splints, red litmus paper, sodium, calcium, powdered magnesium.

Demonstration. 2 ring stands, 1 ring, 2 clamps, wire gauze, 2 Bunsen burners, fish-tail burner, Florence flask (500 cc.), 2 delivery tubes, Pyrex glass tube fitted with 2 one-hole rubber stoppers, collecting bottle, water trough; powdered iron, wood splints, red litmus paper. (See diagram, page 21.)

Unit 7. The Electrolysis of Water (See diagrams, pages 24-25.)

Apparatus. Hoffman apparatus or ring stand with 2 clamps, test tubes, beaker, platinum wire electrodes; copper wire, storage battery or 4 dry cells in series.

Materials. Dilute sulfuric acid or sodium fluoride, wood splints.

Unit 8. Efflorescence and Deliquescence (See diagrams, page 28.)

Apparatus. 2 horn pan or beam balances, 2 watch glasses, 4 test tubes, test tube holder, Bunsen burner.

Materials. Sodium sulfate crystals, sand, sodium hydroxide (sticks), crystals of copper sulfate, sodium carbonate, barium chloride, and potassium nitrate.

Unit 9. Valence, Formulas, and Equations (Neither apparatus nor materials needed.)

Unit 10. Quantitative Chemistry

Experiment A. Balance and weights, ring stand and ring, pipe-stem triangle, Bunsen burner, crucible; crystallized barium chloride.

Experiment B. Balance and weights, thermometer, barometer, Bunsen burner, ring stand, clamp, water trough, hard glass test tube fitted with a one-hole rubber stopper, delivery tube, 2½ liter acid bottle, graduated cylinder; potassium chlorate, manganese dioxide.

Experiment C. Battery jar, 50 cc. eudiometer tube, thermometer, barometer, balance and weights; magnesium ribbon, cotton thread, concentrated hydrochloric acid. (See diagram, page 45.)

Unit 11. Air

Apparatus. Ring stand with ring, wire gauze, small porcelain crucible, flask fitted with solid rubber stopper, Bunsen burner, battery jar, graduated cylinder, 2 watch glasses, rubber bands, iron wire. (See diagram, page 48.)

Materials. Yellow or red phosphorus, wood splints, sodium hydroxide (sticks), limewater.

Unit 12. Chlorine (See diagram, page 53.)

Apparatus. Ring stand with ring and clamp, wire gauze, Bunsen burner, flask fitted with two-hole rubber stopper, thistle tube, delivery tube, connecting tubes, 9 collecting bottles, 4 two-hole rubber stoppers to fit collecting bottles, 6 glass plates, test tubes, hydrogen generator (see diagram, page 15), combustion spoon.

Materials. Concentrated hydrochloric acid, manganese dioxide, sodium, powdered antimony, paraffin candle or taper, colored calico, hypo solution, blue and red litmus paper.

Unit 13. Bromine (See diagram, page 58.)

Apparatus. Ring stand and clamp, Bunsen burner, forceps, test tube fitted with one-hole rubber stopper, delivery tube, collecting tube, 3 test tubes.

Materials. Sodium bromide, manganese dioxide, sulfuric acid (2:1), carbon disulfide (or carbon tetrachloride), chlorine water, calcium.

Demonstration. Beaker; liquid bromine, powdered antimony.

Unit 14. Iodine and Fluorine (See diagram, page 62.)

Apparatus. Ring stand with ring and clamp, Bunsen burner, wire gauze, evaporating dish, funnel, cotton plug, beaker, glass plate, 7 test tubes, lead dish, forceps, file or stylus.

Materials. Potassium iodide, manganese dioxide, sulfuric acid (2:1), alcohol, carbon tetrachloride, carbon disulfide, starch solution, chlorine water, iodine, powdered antimony, paraffin, hydrofluoric acid.

Unit 15. Hydrogen Chloride and Hydrochloric Acid (See diagram, page 66.)

Apparatus. Ring stand with ring and clamp, wire gauze, flask fitted with two-hole rubber stopper, thistle tube, delivery tube, 5 collecting bottles, connecting tube, 4 glass plates, cardboard, 5 test tubes.

Materials. Sodium chloride, sulfuric acid (2:1), concentrated ammonium hydroxide, magnesium ribbon, wood splints, blue litmus paper, zinc, silver nitrate solution, dilute nitric acid.

Unit 16. Ionization and the Electron Theory (See diagram, page 71.)

Apparatus. 2 glass plates, test tubes, stirring rod, beakers, 2 platinum or carbon electrodes, storage battery or 4 dry cells or house current, electric light bulbs.

Materials. Tartaric acid, slaked lime, red and blue litmus paper, copper bromide, methyl (or ethyl) alcohol, anhydrous copper sulfate, distilled water, 10% solutions of the substances listed in the table in Unit 16.

Unit 17. Oxidation and Reduction

Apparatus. Bunsen burner, forceps, hard glass test tube fitted with one-hole rubber stopper, delivery tube, test tube holder, test tubes, test tube rack.

Materials. Powdered magnesium, magnesium ribbon, dry ice, copper oxide, powdered charcoal, ferrous chloride solution, concentrated hydrochloric acid, hydrogen peroxide, potassium, ferrocyanide solution or potassium sulfo cyanate solution, iron filings, potassium ferricyanide solution, ferric chloride solution, lime-water.

Unit 18. Acids, Bases, and Salts

Apparatus. 10 test tubes, glass plate, test tube rack, evaporating dish, wire gauze, ring stand and ring, Bunsen burner, 2 burettes, burette stand, 2 burette clamps, 2 Erlenmeyer flasks, 2 stirring rods.

Materials. Red and blue litmus paper; dilute solutions of hydrochloric, nitric, sulfuric, and acetic acids; dilute solutions of sodium hydroxide, potassium hydroxide, and ammonium hydroxide; saturated solution of calcium hydroxide, dilute (1:10 approx.) hydrochloric acid, molar solution of sodium hydroxide, phenolphthalein solution (1 g. in 100 cc. alcohol).

Unit 19. Solutions, Suspensions, and Colloidal Suspensions

Apparatus. 4 wide mouth bottles, 5 test tubes, filter paper, funnel, ring stand and ring, test tube holder, Bunsen burner.

Materials. Salt, sand, starch, potassium permanganate, dilute iodine solution.

Unit 20. Aids to Solution

Apparatus. Bunsen burner, mortar and pestle, 5 test tubes, test tube holder, stirring rod.

Materials. Copper sulfate crystals of uniform size, potassium nitrate, ammonium hydroxide solution, red litmus paper.

Unit 21. Distillation of Water (See diagrams, pages 96-97.)

Apparatus. Ring stand with ring and clamp, Bunsen burner, wire gauze, flask fitted with one-hole rubber stopper, delivery tube, 8 test tubes, test tube rack, beaker.

Materials. Copper sulfate, sodium chloride, phenolphthalein solution, dilute ammonium hydroxide.

Demonstration. Distillation apparatus with Liebig condenser.

Unit 22. Hydrolysis

Apparatus. Test tube rack, 10 test tubes, stirring rod.

Materials. Red and blue litmus paper, solutions of sodium carbonate, copper sulfate, sodium sulfate, sodium chloride, aluminum sulfate, potassium nitrate, ferric chloride, sodium phosphate, zinc sulfate, sodium tetraborate (borax).

Unit 23. Why Reactions Go to Completion

Apparatus. 7 test tubes, test tube holder, Bunsen burner, beaker.

Materials. Red and blue litmus paper; sodium chloride; ammonium chloride; solutions of calcium chloride, silver nitrate, sodium sulfate, barium chloride, sodium nitrate, sodium chloride, potassium nitrate; sulfuric acid (2:1), dilute sodium hydroxide, dilute hydrochloric acid.

Unit 24. Ammonia (See diagram, page 110.)

Apparatus. Bunsen burner, ring stand, clamp, test tube holder, 6 test tubes, collecting bottle, one-hole rubber stopper to fit test tube, delivery tube, beaker (or pan).

Materials. Wood splints, red and blue litmus paper, slaked lime, ammonium chloride, ammonium sulfate, sodium hydroxide solution, concentrated hydrochloric acid.

Unit 25. Nitric Acid (See diagram, page 114.)

Apparatus. Funnel, retort, tripod, Bunsen burner, ring stand and clamp, wire gauze, beaker, test tube holder, 6 test tubes.

Materials. Sodium nitrate, concentrated sulfuric acid, red and blue litmus paper, wood splints, white of cooked egg, copper, ferrous sulfate (freshly prepared solution).

Unit 26. Sulfur

Apparatus. 4 test tubes, test tube holder, Bunsen burner, watch glass, magnifying lens, beaker.

Materials. Powdered roll sulfur, filter paper, copper foil or wire, silver coin, carbon disulfide.

Demonstration. Asbestos mat, Bunsen burner; powdered sulfur, powdered zinc.

Unit 27. Hydrogen Sulfide (See diagram, page 123.)

Apparatus. Ring stand and clamp, side-arm test tube, delivery tube, collecting tube, Bunsen burner, 6 test tubes, test tube rack, blowpipe or jet tube.

Materials. Ferrous sulfide, dilute hydrochloric or sulfuric acid, blue and red litmus, silver coin, filter paper, lead acetate solution, solutions of copper sulfate, cadmium nitrate, antimony chloride, and zinc nitrate.

Unit 28. Sulfur Dioxide (See diagram, page 126.)

Apparatus. Ring stand with ring and clamp, Bunsen burner, wire gauze, flask fitted with two-hole rubber stopper, thistle tube, delivery tube, 5 collecting bottles, connecting tubes, combustion spoon. (A dropping funnel, if available, is suggested in place of the thistle tube.)

Materials. Roll sulfur, wood splints, sodium sulfite, dilute sulfuric acid, red and blue litmus paper, dilute potassium permanganate solution, colored flower or red apple skin.

Unit 29. Sulfuric Acid

Apparatus. 8 test tubes, beaker, test tube holder, Bunsen burner, test tube rack.

Materials. Concentrated sulfuric acid, wood splints, cane sugar, granulated zinc, copper, dilute sulfuric acid.

Demonstration. Apparatus in diagram on page 130; sulfur, platinum wire, test tube, red and blue litmus paper, barium chloride solution, concentrated hydrochloric acid.

Unit 30. Carbon

Apparatus. 3 test tubes, cork to fit test tube, funnel, Bunsen burner, test tube holder, blowpipe.

Materials. Powdered charcoal, dilute hydrogen sulfide solution, powdered boneblack, solution of brown sugar, litharge (PbO), filter paper, charcoal blocks.

Demonstration. 8 test tubes, test tube rack; powdered charcoal, lamp black, carbon disulfide, carbon tetrachloride, solutions

of sodium hydroxide, ammonium hydroxide, nitric acid, hydrochloric acid.

Unit 31. Destructive Distillation (See diagram, page 139.)

Apparatus. Bunsen burner, ring stand, clamp, 2 hard glass test tubes, one-hole rubber stopper to fit test tube, delivery tube, 2 wide mouth bottles fitted with two-hole rubber stoppers, 2 jet tubes.

Materials. Soft coal, wood splints, litmus paper, lead acetate paper.

Unit 32. Carbon Dioxide (See diagram, page 142.)

Apparatus. 5 wide mouth bottles, two-hole rubber stopper to fit wide mouth bottle, thistle tube, delivery tube, water trough, 3 glass plates, 2 test tubes, right angle glass tube.

Materials. Marble chips, dilute hydrochloric acid, limewater, candle, wood splints, red and blue litmus paper.

Unit 33. Uses of Carbon Dioxide (See diagram, page 146.)

Apparatus. Ring stand, clamp, test tube fitted with one-hole rubber stopper, delivery tube, collecting tube, 6 test tubes, wide mouth bottle fitted with one-hole stopper, glass nozzle.

Materials. Sodium bicarbonate, dilute hydrochloric acid, dilute sulfuric acid, cream of tartar, red and blue litmus paper, limewater, commercial baking powders.

Unit 34. Metallurgy (See diagram, page 150.)

Apparatus. Bunsen burner, blowpipe, crucible, forceps, beaker, 2 carbon electrodes, 2 dry cells or battery, ring stand with ring, pipe-stem triangle.

Materials. Charcoal block, litharge, lead sulfide, copper sulfate solution.

Unit 35. Relative Activity of Metals

Apparatus. Test tube rack, 12 test tubes, Bunsen burner.

Materials. Zinc, copper, wood splints, solutions of hydrochloric acid, lead nitrate, silver nitrate, mercurous nitrate, copper sulfate, aluminum sulfate.

Unit 36. Tests for Metals, Acid Radicals, and Some Common Gases

Apparatus. Bunsen burner, blowpipe, thin glass stirring rods (or if available platinum wires sealed in glass rods), blue cobalt glass.

Materials. Charcoal block; borax; cobalt nitrate solution; vials or bottles equipped with cork stoppers each of which has inserted in it a platinum or nichrome wire and containing compounds of sodium, potassium, calcium, lithium, barium, and a mixture of a sodium and a potassium compound; compounds of zinc, aluminum, magnesium, cobalt, chromium, iron, nickel, and a hydrogen sulfide generator (or a freshly prepared H_2S or $(\text{NH}_4)_2\text{S}$ solution).

Unit 37. Identification of an Unknown Salt

All the apparatus and materials of Unit 36. In addition, a clean platinum or nichrome wire one end sealed in a glass rod, red and blue litmus paper, carbon disulfide, solutions of sodium hydroxide, nitric acid, silver nitrate, dilute hydrochloric acid, limewater, chlorine water, lead acetate, barium chloride, freshly prepared ferrous sulfate, and concentrated sulfuric acid.

Unit 38. Aluminum and Some of Its Compounds

Apparatus. 2 test tubes, 2 glass cylinders or wide mouth bottles, evaporating dish, tripod, pipestem triangle, Bunsen burner, wire gauze, forceps, stirring rod.

Materials. Aluminum sulfate solution, ammonium hydroxide, fine clay, cotton cloth, logwood solution.

Demonstration. Ring stand, ring, pan; sand, magnesium ribbon, filter paper, aluminum, and ferric oxide (powdered). (See diagram, page 163.)

Unit 39. Hard Water

Apparatus. 6 test tubes, carbon dioxide generator (see diagram, page 142), test tube holder, Bunsen burner, one-hole stopper to fit test tube, delivery tube, funnel.

Materials. Distilled water, soap solution, limewater, filter paper, calcium sulfate solution, sodium carbonate solution, marble chips, dilute hydrochloric acid.

Unit 40. Calcium Compounds

Apparatus. Tripod, Bunsen burner, wire gauze, forceps, 2 test tubes, evaporating dish, beaker, stirring rod.

Materials. Marble chips, red and blue litmus paper, plaster of Paris, kerosene, 2 match boxes, coin, mortar, Portland cement, sand.

Unit 41. Hydrocarbons and Other Organic Compounds (See diagrams, pages 174-175.)

Apparatus. Ring stand, clamp, hard glass tube fitted with one-hole rubber stopper, delivery tube, water trough, collecting tube, Bunsen burner, 2 wide mouth bottles, cork to fit wide mouth bottle, forceps, glass stirring rod, 2 watch glasses, 2 glass plates.

Materials. Sodium acetate, soda-lime, dilute bromine water, calcium carbide, red and blue litmus paper, gasoline, kerosene, wood splints.

Unit 42. Soap

Apparatus. 5 test tubes, evaporating dish, tripod, wire gauze, stirring rod, Bunsen burner, beaker, spatula or glass plate, funnel, filter stand.

Materials. Olive oil or beef fat or cottonseed oil, 30% solution of sodium hydroxide, alcohol, saturated salt solution, soap solution, red and blue litmus paper, kerosene, lamp black, filter paper, calcium sulfate solution, distilled water.

Unit 43. Paints and Lacquers

Apparatus. Beakers (2), test tubes, solid rubber stopper for test tube, thermometer (0° - 150°C), metal dish, blowpipe, mortar and pestle.

Materials. Wood splints or applicators, cottonseed oil, raw linseed oil, boiled linseed oil, motor oil, turpentine, Japan drier, white "lead" paint, absorbent cotton, wood surface to be painted,

white "zinc" paint, hydrogen sulfide generator (or ammonium sulfide solution), carbon tetrachloride, charcoal block, dyes (such as Gentian Violet, methylene blue, or Malachite green), collodion, olive oil.

Unit 44. Food Ingredients

Apparatus. Bunsen burner, test tube rack, 8 test tubes, test tube holder, porcelain crucible cover, tripod, pipestem triangle.

Materials. Sodium carbonate, red and blue litmus paper, Benedict's (or Fehling's) solution, tincture of iodine, dilute hydrochloric acid, concentrated sulfuric acid, concentrated nitric acid, concentrated ammonium hydroxide, glucose solution, olive oil, gasoline (or benzol), starch, potato, bread, apple, cane sugar, ground peanuts, hard boiled egg, meat.

Unit 45. Consumer Tests of Certain Foods

Apparatus. Beakers (3), evaporating dish, stirring rod, watch glasses (2), separatory funnel, test tubes, Bunsen burner, combustion spoon, 8 oz. bottle, pan or trough, small vial, funnel.

Materials. Granulated zinc, dilute hydrochloric acid, concentrated hydrochloric acid, turmeric paper, ammonium hydroxide, carbon tetrachloride, ferric chloride, cloth for filtering, white wool cloth, litmus paper (red and blue), tincture of iodine (or iodine solution), filter paper, sodium carbonate, sodium chloride, colorless gasoline, various foods to be tested (mentioned in the unit).

Unit 46. Textiles

Apparatus. Microscope (5, under each of which is a slide showing a fiber), Bunsen burner, test tube rack, 6 test tubes, test tube holder, forceps, stirring rod.

Materials. Prepared slides of cotton, wool, silk, linen, and rayon fibers; samples of cloth of each of these textiles, 20% solution of sodium hydroxide, dilute ammonium hydroxide, concentrated hydrochloric acid, concentrated sulfuric acid.

Unit 47. Removal of Spots and Stains

Apparatus. Test tube rack, 4 test tubes.

Materials. Cheesecloth, stained pieces of cloth (see Note in

Unit 46), white blotting paper, soap, carbon tetrachloride, turpentine, benzol, alcohol, ether, ammonium oxalate solution, dilute hydrochloric acid, sodium hypochlorite solution, vinegar, lemon juice, tincture of soap, milk, oxalic acid.

Unit 48. Photography

Apparatus. 4 test tubes, test tube rack.

Materials. Dilute solutions of silver nitrate and potassium bromide (about 15 grams per liter), developer solution freshly made from the commercial M-Q developer tubes, 30% solution of sodium thiosulfate (Hypo), black paper, Velox print paper. (See also directions in Unit 48.)

Unit 49. Colloids

Apparatus. Stereopticon or similar projector, test tubes, mortar and pestle, Bunsen burner, beakers (2), funnel.

Materials. Coarse starch, tincture of iodine (or iodine solution), filter paper, distilled water, sugar, sulfur, alcohol, sodium thiosulfate (Hypo), dilute hydrochloric acid, ferric chloride solution, kerosene, liquid soap, Jello (or similar gelatine preparation), lampblack, sodium chloride solution, silver nitrate solution, gelatin solution (1 percent), dye (such as methylene blue), powdered charcoal, copper sulfate solution.

Unit 50. Consumer Tests of Commonly Used Materials

Apparatus. Test tubes, Bunsen burner, beakers (2), blowpipe, evaporating dish, platinum or nichrome wire, funnel.

Materials. Dilute hydrochloric acid, filter paper, new microscope slides, smooth nickel coin, iodine solution, turmeric paper, ammonium hydroxide, concentrated hydrochloric acid, litmus paper (red and blue), ammonium sulfide solution (or a hydrogen sulfide generator), cobalt nitrate, charcoal block, white wool, ammonium oxalate solution, nitric acid, potassium sulfocyanate, strip of copper, stannous chloride, potassium chromate solution, sodium hydroxide solution, lead acetate solution, hydrogen peroxide solution, sodium (or barium or calcium) sulfide; materials

listed in the unit for the composition of a typical face powder, a typical cold cream, a deodorant solution, or a deodorant powder.

Unit 51. Plants and Soils

Apparatus. A. Small pan, balance and weights, thermometer (0° - 150°), crucible, Bunsen burner, tripod, wire gauze, flask, 150 cc. beakers (2), stirring rod, wash bottle, funnel, test tube, cobalt glass, platinum *or* nichrome wire, hard glass (Pyrex) test tube, electric hot plate.

B. 2 watch glasses, 5 gal. bottle, flower pot, glass tubing, liter bottle, wire support for plant.

Materials. A. Finely cut leaves and stems, distilled water, filter paper, red litmus paper, concentrated nitric acid, concentrated hydrochloric acid, dilute ammonium hydroxide, potassium thiocyanate solution, ammonium oxalate solution, sodium phosphate solution, concentrated ammonium hydroxide, silver nitrate solution, barium chloride solution, ammonium molybdate solution, soda lime, dilute iodine solution, potato (corn or bean), alcohol.

B. Sample of soil, red and blue litmus paper, distilled water, sand, seeds, nitrazine paper, mono-potassium phosphate, calcium nitrate, magnesium sulfate, ammonium sulfate, boric acid, copper sulfate, manganese sulfate, zinc sulfate, ferrous sulfate (or ferric chloride).

Unit 52. Review of Equations (Neither apparatus nor materials needed.)

Unit 53. Chemical Arithmetic (Neither apparatus nor materials needed.)

Unit 54. Biography (Neither apparatus nor materials needed.)

Unit 55. Table of Common Substances and Alloys (Neither apparatus nor materials needed.)

Unit 56. Review of Important Terms and Laws (Neither apparatus nor materials needed.)

Individual Apparatus Kit
for
Secondary School Chemistry Course

- | | |
|-----------------------------------|----------------------------------|
| 2 Beakers, "Pyrex," 250 cc. | Labels, small box, 1" square |
| 1 Blow pipe, brass, 8" | Litmus paper, both red and |
| 5 Bottles, wide mouthed, 8 oz. | blue, 20 pieces each, 1"x1¼" |
| 1 Bunsen burner with tubing | 1 Nichrome wire, 3" long |
| 1 Brush, test tube, fan tip | 1 Ring stand with 2" and 3" ring |
| 1 Calcium chloride tube | Rubber stoppers |
| 1 Charcoal block | 1 one-hole stopper No. 1 |
| 1 Clamp (burette clamp type) | 1 two-hole stopper No. 6 or |
| 1 Clamp holder | No. 7 |
| 1 Cobalt glass | 1 solid stopper No. 4 |
| 1 Deflagrating (combustion) | 1 two-hole stopper No. 4 |
| spoon | 2 ft. Rubber tubing, ¼" inside |
| 1. Evaporating dish, Sillimanite, | diameter |
| No. 0 | 50 Splints, wood |
| 1 Funnel, glass, 65 mm. | 12 Test tubes, soft, 6"x¾" |
| 1 File, triangular | 2 Test tubes, Pyrex, 6"x¾" |
| Filter paper, small pkg., | 1 Test tube holder |
| 5" diameter | 1 Test tube rack |
| 1 Forceps, laboratory, 5" | 1 Thistle tube |
| 1 Flask, Florence, "Pyrex," | 1 Triangle, pipe stemmed, 2" |
| 250 cc. | 1 Tripod |
| 5 ft. Glass tubing, ¼" external | 1 Tongs (pair) |
| diameter | 1 Watch glass, 3½" |
| 2 ft. Glass rod, ⅜" diameter | 1 Wing tip for Bunsen burner |
| 4 Glass plates, 3" squares | 1 Wire gauze square, asbestos |
| 1 Graduate, cylindrical, 50 cc. | centers, 5" square |

Apparatus for Demonstrations

This list consists of apparatus not already included in students laboratory work. Unless otherwise marked, the quantity called for is *one*.

Acetylene burner	Conductivity of solution apparatus
Alcohol burner	
Alpha ray track apparatus*	2 Condensers, Liebig
2 Asbestos boards, (6"x6")	2 Condenser clamps
Atomic hydrogen torch	Cork borers 1-6, (1 set)
Atomic Numbers, Table of	Cork press
Balance, 1 analytical, 2 horn pan	2 Crucibles, clay, (4"x2 ³ / ₈ ")
Balance, triple beam	2 Crucibles, porcelain, 10 cc., with cover
Balance weights	2 Crystallizing dishes, 6"
Bar magnet	Dewar flask
Barometer, aneroid	Dessicator, 5"
Battery, lead storage	Dialyzer apparatus
2 Battery jars, (10 cm. x 12 cm.)	Diffusion thimbles
Blowpipe, oxyacetylene	12 Dry cells
Bellows, foot	Drying oven*
Beakers, "Pyrex,"	Electrolysis of water apparatus
250 cc. (3)	Eudiometer, 50 cc.
400 cc. (2)	Fire extinguisher, portable
1000 cc. (1)	2 Flasks, side arm, 250 cc.
Blast lamp	Funnel, separatory, 60 cc.
4 doz. Bottles, reagent	Gas mantle, cerium oxide
Brownian movement apparatus*	Generator, Kipp*
2 Burettes, 50 cc.	Glass cutter
Calorimeter, bomb	Graduates, cylindrical,
Cathode ray tube*	100 cc. (1)
Calcium chloride tube, 6"	250 cc. (2)
1 doz. Candles, paraffin	500 cc. (1)
Carbon plates	Hydrometer jar
Carbon rods	Induction cell
Combustion tube	Lead dish
	2 Lamp chimneys

*Optional

Microscopes, compound , 10 slides	Safety lamp, miner's
6 Medicine droppers	Spectroscope
Mortar and pestle, 6" (iron)	Still
Periodic Table of the Elements	Shears for cutting metal, 12"
Photoelectric cell	2 Test tube racks
Pyrometer	2 Thermometers, Centigrade
2 Photographic trays	—10° to 110° C
3 Retorts, glass, 250 cc., ground	2 Thermometers, Fahrenheit
glass stopper	—10° to 220° F

Chemicals Required for Laboratory Work*

It is assumed that chemicals are to be purchased for a complete laboratory course in secondary school chemistry to meet the needs of a class of twenty-four students working individually. The following list includes amounts of items that will suffice for at least one year. In many instances, amounts are listed that will last for many years. These latter items, however, cannot be ordered in smaller amounts economically.

1 lb. Acetone	1 lb. Ammonium sulfate, C. P.
2 lb. Acid, acetic, glacial	1 lb. Ammonium sulfide, C. P.
2 oz. Acid, citric	4 oz. Antimony, powder
4 oz. Acid, formic	8 oz. Antimony chloride, C. P.
30 lb. Acid, hydrochloric, C. P., Sp. gr. 1.18	1 lb. Asbestos, fibre, shreds, paper
10 lb. Acid, nitric, C. P., Sp. gr. 1.42	3 lb. Barium chloride, C. P., crystals
4 oz. Acid, oxalic, crystals, C. P.	2 qt. Benzol
15 lb. Acid, sulfuric, Sp. gr. 1.84	1 lb. Bleaching powder
2 oz. Acid, tartaric	3 lb. Boneblack
1 lb. Alcohol, amyl	2 lb. Brass, sheet, 18 gage
1 gal. Alcohol, ethyl (denatured)	2 oz. Bromine
1 pt. Alcohol, ethyl, U. S. P.	$\frac{1}{2}$ lb. Calcite crystals
2 pt. Alcohol, methyl	8 oz. Calcium metal
2 oz. Alizarine paste	1 lb. Calcium carbide, tech.
1 lb. Aluminum powder (very fine)	10 lb. Calcium carbonate, marble chips
10 lb. Aluminum sheet, 18 gage	2 lb. Calcium chloride, granular, for drying tubes
1 lb. Aluminum sulfate, C. P.	3 lb. Calcium hydroxide
4 lb. Ammonium chloride, C. P.	8 oz. Calcium nitrate, C. P.
20 lb. Ammonium hydroxide	10 lb. Calcium oxide, tech. lump
2 oz. Ammonium oxalate, crystal, C. P.	8 oz. Calcium phosphate (mono calcium)

*In making these lists, it has been our intention to include those chemicals and equipment which will permit all experiments and demonstrations suitable for secondary school chemistry students. It has been assumed that each teacher will eliminate those items not needed in his program.

10 lb. Calcium sulfate, plaster of Paris, fine	8 oz. Fehling's solutions, A and B
3 lb. Carbon disulfide, C. P.	2 oz. Fluorescein
1 lb. Carbon tetrachloride	1 qt. Gasoline
4 oz. Casein	4 oz. Gelatine
2 lb. Cement, Portland	1 lb. Glucose
1 lb. Chalk, precipitated	1 lb. Glycerine, U. S. P.
50 sticks Charcoal, blocks	1 lb. Gypsum
1 lb. Charcoal, wood, powdered	8 oz. Hydrochinone
1 lb. Chloroform, U. S. P.	1 lb. Hydrofluoric acid, 48% C. P.
1 lb. Chrome alum, C. P.	2 lb. Hydrogen peroxide, 3% U. S. P.
2 yd. Cloth, calico, bleachable color, for bleaching with chlorine	4 oz. Iodine, crystal, C. P.
2 yd. Cloth cotton, bleached, fine goods	1 lb. Iron nails 2"
4 oz. Cobalt nitrate, crystals, C. P.	1 lb. Iron and ammonium citrate, C. P.
1 lb. Copper bromide, C. P.	1 lb. Iron chloride (ferric) C. P.
1 lb. Copper foil, $\frac{1}{100}$ " thick	1 lb. Iron filings, fine, clean
10 sq. in. Copper gauze, 80 meshes to inch	1 lb. Iron oxide (jeweler's rouge)
2 lb. Copper rivets, $\frac{1}{2}$ "	1 lb. Iron sulfate (ferrous) C. P.
5 lb. Copper sheets, $\frac{1}{64}$ " thick	1 lb. Iron sulfide (stick form)
1 lb. Copper wire, spool No. 18	1 lb. Iron sheet, 18 gage
1 lb. Copper wire, spool No. 24	1 lb. Iron wire, spool No. 16
1 lb. Copper wire, spool No. 30	1 lb. Iron wire, spool No. 22
1 lb. Copper oxide, C. P., powder	1 pt. Kerosene
1 lb. Copper oxide, C. P., wire form	1 lb. Lampblack
1 lb. Copper sulfate, anhydrous, tech.	2 lb. Lead acetate, C. P.
5 lb. Copper sulfate, C. P., crystals	2 lb. Lead nitrate, C. P.
2 lb. Cotton, absorbent	1 lb. Lead oxide (litharge)
2 lb. Cotton seed oil	1 lb. Lead sulfide
2 oz. Dextrine	2 lb. Lead sheet
1 lb. Ether, U. S. P.	1 lb. Lead shot No. 10
	1 lb. Lithium nitrate
	4 oz. Litmus cubes
	1 quire Litmus paper, red and blue (each)
	8 oz. Logwood, ground, tech.
	8 oz. Magnesium powder, U. S. P.

8 oz. Magnesium ribbon, U. S. P.	2 lb. Salt, rock
1 lb. Magnesium sulfate, crystals	5 lb. Sand
1 lb. Manganese chloride, C. P.	8 oz. Silver nitrate, U. S. P.
5 lb. Manganese dioxide, fine powder	4 oz. Soap, castile, powder
2 lb. Mercury, redistilled, 5 lb. jug	8 oz. Soda-lime
4 oz. Mercuric nitrate C. P.	4 oz. Sodium
4 oz. Mercurous nitrate, C. P.	1 lb. Sodium Acetate
1 lb. Mercuric oxide, C. P.	1 lb. Sodium alum
1 lb. Molasses	3 lb. Sodium bicarbonate
1 lb. Nichrome wire No. 18	1 lb. Sodium bisulfite
1 pt. Olive oil	1 lb. Sodium bromide
2 lb. Paraffin	2 lb. Sodium carbonate (washing soda)
20 gm. Phenolphthalein	1 lb. Sodium carbonate, pure, dry
4 oz. Phosphorous, red	10 lb. Sodium chloride (table salt)
4 oz. Phosphorous, yellow	1 lb. Sodium chromate, C. P.
1 gm. Platinum wire	5 lb. Sodium hydroxide, C. P., sticks
2 lb. Potassium aluminum sulfate, C. P.	2 qt. Sodium hypochlorite solution
1 lb. Potassium acid tartrate	1 lb. Sodium iodide, U. S. P.
1 lb. Potassium bromide	1 lb. Sodium fluoride
4 lb. Potassium chlorate, crystals, C. P.	2 lb. Sodium nitrate
4 oz. Potassium chloride, C. P.	4 oz. Sodium peroxide
8 oz. Potassium chromate, C. P.	1 lb. Sodium phosphate
2 lb. Potassium dichromate, C. P.	1 lb. Sodium potassium tartrate
1 lb. Potassium ferricyanide, C. P.	1 lb. Sodium silicate
8 oz. Potassium ferrocyanide, C. P.	2 lb. Sodium sulfate, crystals
1 lb. Potassium hydroxide, C. P.	2 lb. Sodium sulfite, pure, dry
1 lb. Potassium iodide, U. S. P.	2 lb. Sodium thiosulfate (hypo)
2 lb. Potassium nitrate, crystal, C. P.	2 lb. Sodium tetraborate (borax)
1 lb. Potassium permanganate, C. P.	2 lb. Starch, corn
4 oz. Potassium sulfocyanate, C. P.	1 lb. Strontium nitrate, C. P.
	2 lb. Sucrose (sugar)
	1 lb. Sugar, brown

1 lb. Sulfur, flowers	2 pkg. Wax tapers
4 lb. Sulfur, roll	8 oz. Wool, glass
8 oz. Tartaric acid, C. P., crystals	1 lb. Wool, steel
8 oz. Tin foil $\frac{1}{100}$ "	5 lb. Zinc sheet, 18 gage
8 oz. Tincture of iodine	8 oz. Zinc dust
1 lb. Tincture of green soap	10 lb. Zinc, mossy
8 lb. Tin, granulated	1 lb. Zinc chloride, C. P.
8 lb. Tin, sticks	1 lb. Zinc nitrate, C. P.
2 lb. Turpentine	1 lb. Zinc sulfate, C. P.

Chemicals Required for Demonstrations

The chemicals listed below are those not already included in the list of chemicals required for laboratory work. In each case one-fourth pound should be ordered, unless otherwise stated.

Acetylene (small cylinder)	Carnallite (sample)
Agar agar	Cassiterite (sample)
Alabaster (sample)	Cast iron (sample)
Alum, common	Celanese (sample)
Amethyst (sample)	Cellophane (sample)
Alundum (sample)	Celluloid (sample)
Ammonium acetate	1 lb. Charcoal, activated
Ammonium nitrate	1 lb. Charcoal, animal
Ammonium nitrite	Chromic acid
Amyl alcohol	Chromite (sample)
Aniline	Cinnabar (sample)
1 oz. Arsenic	1 lb. Clay, powdered
1 oz. Arsenic chloride	1 oz. Cobalt chloride
Asbestos (sample)	Coke (sample)
Babbitt metal (sample)	Collodion
Bakelite (sample)	Congo Red (sample)
Baking powders (samples)	Coral (sample)
Barium dioxide	1 oz. Coper nitrate
Bauxite (sample)	1 pt. Cottonseed oil
1 pt. Benzine	Cream of tartar (sample)
1 oz. Bichloride of mercury	Cryolite (sample)
1 oz. Bismuth	Cyanamid (sample)
Boric Acid	Dextrose
Cadmium nitrate	1 oz. Diastase
Cadmium sulfide	Duprene (sample)
Calcite (sample)	Duralumin (sample)
Calcium arsenate	Duriron (sample)
Calcium fluoride (sample)	Eosine (sample)
Calcium hypochlorite	Emery (sample)
Calcium superphosphate	Ferrous ammonium citrate
Caliche (sample)	Fibers (samples)
Camphor (sample)	1 lb. Formaldehyde
Carboloy (sample)	Formalin
Carborundum (sample)	Fuller's earth (sample)

Galena (sample)	Opal (sample)
Galvanized iron (sample)	Paris green (sample)
Gall nuts (samples)	Peat (sample)
German silver (sample)	Permalloy (sample)
Glass wool (sample)	1 lb. Permutit
Graphite (sample)	Petroleum, crude (sample)
Gum arabic	Pewter (sample)
Hematite (sample)	Prussian blue (sample)
Hoolamite (sample)	Pyrogallol
Hopcalite (sample)	Quartz (samples)
1 oz. Indigo	Rock phosphate (sample)
Insulin (sample)	Ruby (sample)
Invar (sample)	Saccharine (sample)
1 oz. Iodoform	1 oz. Salicylic acid
Iron, pig, Russia, wrought (samples)	Sapphire (sample)
1 pt. Javelle water	Selenium glass (sample)
Kaolin (sample)	1 lb. Sodium bisulfate
1 oz. Lactic acid	1 oz. Sodium cyanide
1 lb. Lactose	Solder (sample)
1 oz. Levulose	Stainless steel (sample)
Lignin (sample)	Stalactite (sample)
Lignite (sample)	Stalagmite (sample)
Lime (sample)	1 lb. Stearic acid
1 qt. Limewater	Tannic acid
Lithopone (sample)	Tetraethyl lead (sample)
Magnalium (sample)	1 oz. Thermit mixture, can
Mauve (sample)	Thyroxin (sample)
1 oz. Mercurous chloride	Tung oil (sample)
Methylene blue (sample)	Turnbull's blue
Methyl orange (sample)	Vaseline
Metel (sample)	Vitamin C (sample)
Milk of magnesia (sample)	Vitreosil (sample)
Monel metal (sample)	Water glass
Naphtha	Wax
Naphthalene	1 lb. White lead
Nigrosine (sample)	Wood's metal (sample)
Nitrobenzine	Zinc stearate
Onyx (sample)	1 pt. Zonite
	Zymase (sample)

THE NINETY-TWO CHEMICAL ELEMENTS

AT. No.	ELEMENT	SYM- BOL	ATOMIC WEIGHT	AT. No.	ELEMENT	SYM- BOL	ATOMIC WEIGHT
1	Hydrogen	H	1.0080	47	Silver	Ag	107.880
2	Helium	He	4.003	48	Cadmium	Cd	112.41
3	Lithium	Li	6.940	49	Indium	In	114.76
4	Beryllium	Be	9.02	50	Tin	Sn	118.70
5	Boron	B	10.82	51	Antimony	Sb	121.76
6	Carbon	C	12.010	52	Tellurium	Te	127.61
7	Nitrogen	N	14.008	53	Iodine	I	126.92
8	Oxygen	O	16.0000	54	Xenon	Xe	131.3
9	Fluorine	F	19.00	55	Cesium	Cs	132.91
10	Neon	Ne	20.183	56	Barium	Ba	137.36
11	Sodium	Na	22.997	57	Lanthanum	La	138.92
12	Magnesium	Mg	24.32	58	Cerium	Ce	140.13
13	Aluminum	Al	26.97	59	Praseodymium	Pr	140.92
14	Silicon	Si	28.06	60	Neodymium	Nd	144.27
15	Phosphorus	P	30.98	61	Illinium	Il	?
16	Sulfur	S	32.06	62	Samarium	Sm	150.43
17	Chlorine	Cl	35.457	63	Europium	Eu	152.0
18	Argon	A	39.944	64	Gadolinium	Gd	156.9
19	Potassium	K	39.096	65	Terbium	Tb	159.2
20	Calcium	Ca	40.08	66	Dysprosium	Dy	162.46
21	Scandium	Sc	45.10	67	Holmium	Ho	163.5
22	Titanium	Ti	47.90	68	Erbium	Er	167.2
23	Vanadium	V	50.95	69	Thulium	Tm	169.4
24	Chromium	Cr	52.01	70	Ytterbium	Yb	173.04
25	Manganese	Mn	54.93	71	Lutecium	Lu	174.99
26	Iron	Fe	55.85	72	Hafnium	Hf	178.6
27	Cobalt	Co	58.94	73	Tantalum	Ta	180.88
28	Nickel	Ni	58.69	74	Tungsten	W	183.92
29	Copper	Cu	63.57	75	Rhenium	Re	186.31
30	Zinc	Zn	65.38	76	Osmium	Os	190.2
31	Gallium	Ga	69.72	77	Iridium	Ir	193.1
32	Germanium	Ge	72.60	78	Platinum	Pt	195.23
33	Arsenic	As	74.91	79	Gold	Au	197.2
34	Selenium	Se	78.96	80	Mercury	Hg	200.61
35	Bromine	Br	79.916	81	Thallium	Tl	204.39
36	Krypton	Kr	83.7	82	Lead	Pb	207.21
37	Rubidium	Rb	85.48	83	Bismuth	Bi	209.00
38	Strontium	Sr	87.63	84	Polonium	Po	210.0
39	Yttrium	Y	88.92	85	Alabamine	Ab	?
40	Zirconium	Zr	91.22	86	Radon	Rn	222.
41	Columbium	Cb	92.91	87	Virginium	Vi	?
42	Molybdenum	Mo	95.95	88	Radium	Ra	226.05
43	Masurium	Ma	?	89	Actinium	Ac	227.0
44	Ruthenium	Ru	101.7	90	Thorium	Th	232.12
45	Rhodium	Rh	102.91	91	Protoactinium	Pa	231.
46	Palladium	Pd	106.7	92	Uranium	U	238.07

ELECTROCHEMICAL SERIES

1. Lithium	7. Manganese	13. Nickel	19. Bismuth
2. Potassium	8. Zinc	14. Tin	20. Arsenic
3. Sodium	9. Chromium	15. Lead	21. Mercury
4. Calcium	10. Cadmium	16. Hydrogen	22. Silver
5. Magnesium	11. Iron	17. Copper	23. Platinum
6. Aluminum	12. Cobalt	18. Antimony	24. Gold

THE METRIC SYSTEM

1 decimeter (dm.) = 0.1 meter (m.)	1 decigram (dg.) = 0.1 gram (g.)
1 centimeter (cm.) = 0.01 meter (m.)	1 centigram (cg.) = 0.01 gram (g.)
1 millimeter (mm.) = 0.001 meter (m.)	1 milligram (mg.) = 0.001 gram (g.)
1 kilometer (km.) = 1000 meters (m.)	1 kilogram (kg.) = 1000 grams (g.)

Unit Equivalent (approximate)

1 inch (in.)	= 2.54 centimeters (cm.)
1 square inch (sq. in.)	= 6.45 square centimeters (sq. cm.)
1 square meter (sq. m.)	= 10.76 square feet (sq. ft.)
1 cubic inch (cu. in.)	= 16.4 cubic centimeters (cc.)
1 cubic foot (cu. ft.)	= 28.3 liters (l.)
1 cubic foot (cu. ft.)	= 7.48 gallons (gal.)
1 liter (l.) (1000 cc.)	= 61 cubic inches (cu. in.)
1 liter (l.)	= 1.06 quarts (qt.)
1 gallon (gal.)	= 231 cubic inches (cu. in.)
1 cubic meter (cu. m.)	= 1.31 cubic yards (cu. yd.)
1 cubic meter (cu. m.)	= 35.32 cubic feet (cu. ft.)

1 ounce (oz.) Avoirdupois	= 28.35 grams (g.)
1 kilogram (kg.)	= 2.2 pounds (lb.)
1 ton (T.), 2000 lb.	= 907.2 kilograms (kg.)
1 liter of water	= 2.2 pounds (lb.)
1 cubic foot of water	= 62.4 pounds (lb.)
1 gallon of water	= 8.35 pounds (lb.)
1 atmosphere pressure	= 1033 grams per square centimeter

SOME PROPERTIES OF IMPORTANT SOLIDS

SOLID	SPECIFIC GRAVITY (WATER = 1)	HARDNESS	MELTING POINT DEGREES C	TENSILE STRENGTH (LB. PER SQ. IN.)
Aluminum	2.7	2	658.7	30,000—40,000
Copper	8.9	2.5—3	1083	60,000—70,000
Diamond	3.0—3.5	10		
Glass	2.4—2.8	4.5—6.5	1100	2560
Gold	19.3	2.5—3	1063	40,000—50,000
Iron, cast	7.0—7.7	5.0	1250	13,000—33,000
Iron, wrought	7.8—7.9	4.5	1600	50,000
Lead	11.35	1.5	327	2600—3300
Limestone	2.7—2.8	3		
Platinum	21.37	4.3	1755	50,000
Quartz	2.65	7	1600	
Sulfur	2.0—2.1	1.5—2.5	113	
Silver	10.5	2.5—3	960.5	42,000
Steel	7.6—7.8	5—8.5	1350	80,000—330,000
Tin	7.0	1.5	232	4000—5000
Tungsten	18.6—19.1		3400	590,000
Zinc	7.1	2.5	419	7000—30,000

SOME PROPERTIES OF IMPORTANT LIQUIDS

LIQUID	SPECIFIC GRAVITY (WATER = 1)	FREEZING POINT °C	BOILING POINT °C	HEAT OF VAPORIZATION (CALORIES PER GRAM AT BOILING POINT)
Alcohol (grain)	.789	—117.3	78.3	204
Alcohol (wood)	.792	—97.8	64.7	262.8
Carbon tetrachloride	1.595	—23	76.8	46.4
Ether	.714	—116.3	34.5	83.9
Gasoline	.66—.69		70—90	70
Glycerine	1.260	17.9	291	
Kerosene	.80		150—300	
Mercury	13.596	—39	357	65
Turpentine	.90		159	74
Water	1.00	0	100	536

HEAT VALUES OF FUELS

<i>Gases</i>		SMALL CALORIES PER GRAM
Acetylene		11,900
Carbon Monoxide		2,400
Coal gas		11,000
Hydrogen		33,900
Methane		13,000
Producer gas		600
Water gas		4,500
<i>Liquids</i>		
Alcohol (grain)		7,000
Alcohol (wood)		5,300
Benzene		10,000
Fuel oil		
(heavy petroleum residue)		10,300
Gasoline		11,250
Kerosene		11,100
Paraffin oil		9,800
Petroleum oil		10,800 to 11,100
<i>Solids</i>		
Coal (lignite)		3,200 to 6,700
Coal (bituminous)		6,000 to 7,800
Coal (anthracite)		7,000 to 7,600
Paraffin		10,300
Wood (pine)		5,000

BRIEF SOLUBILITY TABLE

All *nitrates*, *chlorates*, and *acetates* are soluble in water.
 All *chlorides*, *bromides*, and *iodides* are soluble except those of Ag, Pb, and Hg.
 All *sulfates* are soluble except PbSO₄, BaSO₄, CaSO₄, HgSO₄, and Ag₂SO₄.
 All *hydroxides* are insoluble except those of Na, K, NH₄, Ca, and Ba.
 All *carbonates*, *silicates*, and *phosphates* are insoluble except those of Na, K, and NH₄.
 All *sulfides* are insoluble except those of Na, K, NH₄, Ca, and Ba.
 All *oxides* are insoluble except those of Na, K, and Ca.

APPROXIMATE ATOMIC WEIGHTS OF THE COMMON ELEMENTS

ELEMENT	SYMBOL	APPROX. ATOMIC WEIGHT	ELEMENT	SYMBOL	APPROX. ATOMIC WEIGHT
Aluminum	Al	27	Lead	Pb	207
Antimony	Sb	121.7	Lithium	Li	7
Arsenic	As	75	Magnesium	Mg	24.3
Barium	Ba	137.4	Manganese	Mn	55
Bismuth	Bi	209	Mercury	Hg	200.6
Boron	B	11	Nickel	Ni	58.7
Bromine	Br	80	Nitrogen	N	14
Cadmium	Cd	112.4	Oxygen	O	16
Calcium	Ca	40	Phosphorus	P	31
Carbon	C	12	Platinum	Pt	195
Chlorine	Cl	35.5	Potassium	K	39
Chromium	Cr	52	Radium	Ra	226
Cobalt	Co	59	Silicon	Si	28
Copper	Cu	63.6	Silver	Ag	108
Fluorine	F	19	Sodium	Na	23
Gold	Au	197	Strontium	Sr	87.6
Hydrogen	H	1	Sulfur	S	32
Iodine	I	127	Tin	Sn	118.7
Iron	Fe	56	Zinc	Zn	65.4

TABLE OF COMMON VALENCES

	MONOVALENT		DIVALENT		TRIVALENT	
METALS	ammonium	NH_4^+	calcium	Ca^{++}	aluminum	Al^{+++}
	mercurous	Hg^+	copper (ic)	Cu^{++}	antimony	Sb^{+++}
	potassium	K^+	magnesium	Mg^{++}	arsenic	As^{+++}
	silver	Ag^+	mercuric	Hg^{++}	chromium	Cr^{+++}
	sodium	Na^+	iron (ferrous)	Fe^{++}	iron (ferric)	Fe^{+++}
NON-METALS			zinc	Zn^{++}		
	bromine	Br^-	oxygen	O^{--}	nitrogen	N^{---}
	chlorine	Cl^-	sulfur	S^{--}	phosphorus	P^{---}
	fluorine	F^-				
RADICALS	iodine	I^-				
	bicarbonate	HCO_3^-	carbonate	CO_3^{--}	phosphate	PO_4^{---}
	chlorate	ClO_3^-	sulfate	SO_4^{--}		
	hydroxide	OH^-	sulfite	SO_3^{--}		
	nitrate	NO_3^-				
	nitrite	NO_2^-				

FIRST AID

Call a physician or rush the student to a hospital or physician in cases of severe injury.

Every chemistry laboratory should have a first aid kit. Near this kit should be posted an emergency chart for chemistry laboratories such as the one obtainable from the Fisher Scientific Company of Pittsburgh, Pennsylvania. Every chemistry laboratory should also have an asbestos blanket readily available for emergencies.

Treatment for the common types of injury in the chemistry laboratory is indicated below.

ACIDS:

On the skin: Wash with plenty of water. Then apply a thick paste of sodium bicarbonate and water.

On the clothing: Neutralize with dilute ammonium hydroxide.

Taken internally: Drink a solution of sodium bicarbonate *or* milk of magnesia *or* limewater.

BASES:

On the skin or clothing: Wash with plenty of water. Then apply dilute lemon juice *or* citric acid solution *or* boric acid solution. Wash again with water.

Taken internally: Drink lemon juice *or* dilute citric acid.

BURNS:

Apply unguentine *or* carron oil (limewater and linseed oil) *or* a thick paste of sodium bicarbonate and water.

CUTS:

Remove any foreign matter with a pair of sterilized forceps. Apply a disinfectant such as hydrogen peroxide *or* a 3.5% solution of tincture of iodine (U. S. P. tincture of iodine is 7%) and then bandage. A solution of ferric chloride can be used to coagulate the blood to stop excessive bleeding.

1.45'

